1

The Basics

Bonding and Molecular Structure

Organic chemistry is a part of our lives at every moment. Organic molecules comprise the tissue of plants as mighty as the redwoods, convey signals from one neuron to the next in animals, store the genetic information of life, and are the food we eat each day. The growth of living things from microbes to elephants rests on organic reactions, and organic reactions provide the energy that drives our muscles and our thought processes.

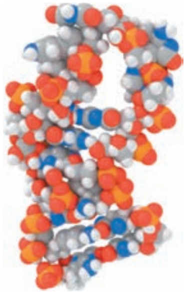
Our lives depend on organic chemistry in many other ways as well. Every article of clothing we wear is a product of organic chemistry, whether the fibers are natural or synthetic. Hardly a minute goes by when we're not using something made of organic molecules, such as a pen, a computer keyboard, a music player, or a cellular phone. We view display screens made of organic liquid crystal arrays. Natural organic polymers comprise wood and the paper we read. Natural and synthetic organic molecules enhance our health. There is not a single aspect of our lives that is not in some way dependent on organic chemistry. But what is organic chemistry?

- **Organic chemistry is the chemistry of compounds that contain the element carbon.**

Clearly, carbon compounds are central to life on this planet. Carbon as an element, however, has its origin elsewhere.

1.1 *We Are Stardust*

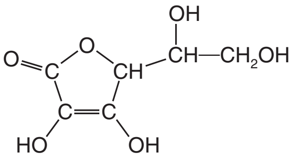
Some 14.5 billion years ago the big bang formed hydrogen and helium, the lightest elements. Further nuclear reactions in stars transmuted these elements into heavier ones, including carbon, nitrogen, oxygen, sulfur, phosphorus, and most others in the periodic table. Massive explosions called supernovae scattered the elements in the universe, and over time heavy elements coalesced to form planets and other celestial bodies. Through processes not understood but about which there continues to be much research, simple molecules formed, eventually including organic molecules that could support life—the nucleic acids that make up DNA and RNA, the amino acids that comprise proteins, carbohydrates such as glucose, and other types of molecules. It is from elegant molecular building blocks like these that the incredible richness of chemistry and life has evolved. So, in the truest sense we living creatures are composed of stardust, and without supernovae not only would there be no organic chemistry, there would be no life.



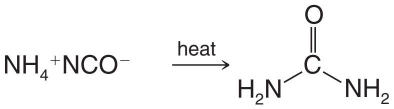
An RNA molecule

1.1A Development of the Science of Organic Chemistry

The science of organic chemistry began to flower with the demise of a nineteenth century theory called vitalism. According to vitalism, organic compounds were only those that came from living organisms, and only living things could synthesize organic compounds through intervention of a vital force. Inorganic compounds were considered those compounds that came from nonliving sources. Friedrich Wöhler, however, discovered in 1828 that an organic compound called urea (a constituent of urine) could be made by evaporating an aqueous solution of the inorganic compound ammonium cyanate. With this discovery, the synthesis of an organic compound, began the evolution of organic chemistry as a scientific discipline.



Vitamin C



$$NH_4^+NCO^- \xrightarrow{\text{heat}} H_2N\overset{O}{\overset{\|}{C}}NH_2$$

Ammonium cyanate **Urea**

Despite the demise of vitalism in science, the word "organic" is still used today by some people to mean "coming from living organisms" as in the terms "organic vitamins" and "organic fertilizers." The commonly used term "organic food" means that the food was grown without the use of synthetic fertilizers and pesticides. An "organic vitamin" means to these people that the vitamin was isolated from a natural source and not synthesized by a chemist. While there are sound arguments to be made against using food contaminated with certain pesticides, while there may be environmental benefits to be obtained from organic farming, and while "natural" vitamins may contain beneficial substances not present in synthetic vitamins, it is impossible to argue that pure "natural" vitamin C, for example, is healthier than pure "synthetic" vitamin C, since the two substances are identical in all respects. In science today, the study of compounds from living organisms is called natural products chemistry.

Vitamin C is found in various citrus fruits.

1.2 *Atomic Structure*

Before we begin our study of the compounds of carbon we need to review some basic but familiar ideas about the chemical elements and their structure.

- The **compounds** we encounter in chemistry are made up of **elements** combined in different proportions. An abridged periodic table of the elements is given in Table 1.1.

TABLE 1.1 An Abridged Periodic Table of the Elements

PERIODIC TABLE OF THE ELEMENTS

Atomic number→ 6
Symbol → C
Name (IUPAC) → Carbon
Atomic mass → 12.011

Chemical Abstracts Service group notation →

IA	IIA											IIIA	IVA	VA	VIA	VIIA	VIIIA
1 H Hydrogen 1.0079																	2 He Helium 4.0026
3 Li Lithium 6.941	4 Be Beryllium 9.0122											5 B Boron 10.811	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.998	10 Ne Neon 20.180
11 Na Sodium 22,990	12 Mg Magnesium 24.305											13 Al Aluminum 26.982	14 Si Silicon 28.086	15 P Phosphorus 30.974	16 S Sulfur 32.065	17 Cl Chlorine 35.453	18 Ar Argon 39.948
19 K Potassium 39.098	20 Ca Calcium 40.078	21 Sc Scandium 44.956	22 Ti Titanium 47.867	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.938	26 Fe Iron 55.845	27 Co Cobalt 58.933	28 Ni Nickel 58.693	29 Cu Copper 63.546	30 Zn Zinc 65.409	31 Ga Gallium 69.723	32 Ge Germanium 72.64	33 As Arsenic 74.922	34 Se Selenium 78.96	35 Br Bromine 79.904	36 Kr Krypton 83.798
37 Rb Rubidium 85.468	38 Sr Strontium 87.62	39 Y Yttrium 88.906	40 Zr Zirconium 91.224	41 Nb Niobium 92.906	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.60	53 I Iodine 126.90	54 Xe Xenon 131.29
55 Cs Cesium 132.91	56 Ba Barium 137.33	57 La Lanthanum 138.91	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.84	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.2	83 Bi Bismuth 208.98	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)
87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium (261)	105 Db Dubnium (262)	106 Sg Seaborgium (266)	107 Bh Bohrium (264)	108 Hs Hassium (277)	109 Mt Meitnerium (268)	110 Uun (281)	111 Uuu (272)	112 Uub (285)		114 Uuq (289)				

(Lanthanide series (58-71) and actinide series (90-103) elements not shown)

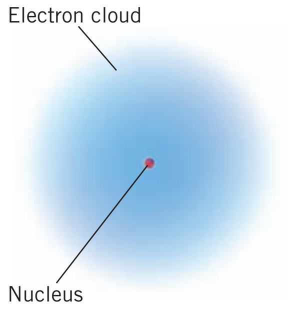


Figure 1.1 An atom is composed of a tiny nucleus containing protons and neutrons and a large surrounding volume containing electrons. The diameter of a typical atom is about 10,000 times the diameter of its nucleus.

- **Elements** are made up of **atoms**. An atom (Fig. 1.1) consists of a dense, positively charged *nucleus* containing **protons** and **neutrons** and a surrounding cloud of **electrons**.

Each proton of the nucleus bears one positive charge; electrons bear one negative charge. Neutrons are electrically neutral; they bear no charge. Protons and neutrons have nearly equal masses (approximately 1 atomic mass unit each) and are about 1800 times as heavy as electrons. Most of the **mass** of an atom, therefore, comes from the mass of the nucleus; the atomic mass contributed by the electrons is negligible. Most of the **volume** of an atom, however, comes from the electrons; the volume of an atom occupied by the electrons is about 10,000 times larger than that of the nucleus.

The elements commonly found in organic molecules are carbon, hydrogen, nitrogen, oxygen, phosphorus, and sulfur, as well as the halogens: fluorine, chlorine, bromine, and iodine.

Each **element** is distinguished by its **atomic number** (*Z*), a **number equal to the number of protons in its nucleus**. Because an atom is electrically neutral, **the atomic number also equals the number of electrons surrounding the nucleus**.

1.2A Isotopes

Before we leave the subject of atomic structure and the periodic table, we need to examine one other observation: **the existence of atoms of the same element that have different masses**.

For example (Table 1.1), the element carbon has six protons in its nucleus giving it an atomic number of 6. Most carbon atoms also have six neutrons in their nuclei, and because each proton and each neutron contributes one atomic mass unit (1 amu) to the mass of the atom, carbon atoms of this kind have a mass number of 12 and are written as ^{12}C.

- **Although all the nuclei of all atoms of the same element will have the same number of protons**, some atoms of the same element **may have different masses** because they have **different numbers of neutrons**. Such atoms are called **isotopes**.

For example, about 1% of the atoms of elemental carbon have nuclei containing 7 neutrons, and thus have a mass number of 13. Such atoms are written ^{13}C. A tiny fraction of carbon atoms have 8 neutrons in their nucleus and a mass number of 14. Unlike atoms of carbon-12 and carbon-13, atoms of carbon-14 are radioactive. The ^{14}C isotope is used in *carbon dating*. The three forms of carbon, ^{12}C, ^{13}C, and ^{14}C, are isotopes of one another.

Most atoms of the element hydrogen have one proton in their nucleus and have no neutron. They have a mass number of 1 and are written ^{1}H. A very small percentage (0.015%) of the hydrogen atoms that occur naturally, however, have one neutron in their nucleus. These atoms, called *deuterium* atoms, have a mass number of 2 and are written ^{2}H. An unstable (and radioactive) isotope of hydrogen, called *tritium* (^{3}H), has two neutrons in its nucleus.

Review Problem 1.1

There are two stable isotopes of nitrogen, ^{14}N and ^{15}N. How many protons and neutrons does each isotope have?

1.2B Valence Electrons

We discuss the electron configurations of atoms in more detail in Section 1.10. For the moment we need only to point out that the electrons that surround the nucleus exist in **shells** of increasing energy and at increasing distances from the nucleus. The most important shell, called the **valence shell**, is the outermost shell because the electrons of this shell are the ones that an atom uses in making chemical bonds with other atoms to form compounds.

- How do we know how many electrons an atom has in its valence shell? We look at the periodic table. The number of electrons in the valence shell (called **valence**

electrons) is equal to the group number of the atom. For example, carbon is in group **IVA** and carbon has *four* valence electrons; oxygen is in group **VIA** and oxygen has *six* valence electrons. The halogens of group **VIIA** all have *seven* electrons.

Review Problem 1.2

How many valence electrons does each of the following atoms have?

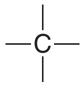
(a) Na **(b)** Cl **(c)** Si **(d)** B **(e)** Ne **(f)** N

1.3 *The Structural Theory of Organic Chemistry*

Between 1858 and 1861, August Kekulé, Archibald Scott Couper, and Alexander M. Butlerov, working independently, laid the basis for one of the most important theories in chemistry: the **structural theory**.

Two central premises are fundamental:

1. The atoms in organic compounds can form a fixed number of bonds using their outermost shell (valence) electrons. Carbon is *tetravalent*; that is, carbon atoms have four valence electrons and can form four bonds. Oxygen is *divalent*, and hydrogen and (usually) the halogens are *monovalent*:



Carbon atoms are tetravalent.

—O—

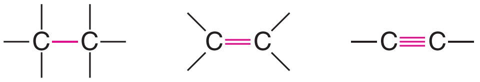
Oxygen atoms are divalent.

H— Cl—

Hydrogen and halogen atoms are monovalent.

2. A carbon atom can use one or more of its valence electrons to form bonds to other carbon atoms:

Carbon–carbon bonds



Single bond **Double bond** **Triple bond**

In his original publication Couper represented these bonds by lines much in the same way that most of the formulas in this book are drawn. In his textbook (published in 1861), Kekulé gave the science of organic chemistry its modern definition: ***a study of the compounds of carbon***.

1.3A Isomers: The Importance of Structural Formulas

The structural theory allowed early organic chemists to begin to solve a fundamental problem that plagued them: the problem of **isomerism**. These chemists frequently found examples of **different compounds that have the same molecular formula**. Such compounds are called **isomers**.

Let us consider an example involving two compounds that have practical uses: acetone, used in nail polish remover and as a paint solvent, and propylene oxide, used with seaweed extracts to make food-grade thickeners and foam stabilizers for beer (among other applications). Both of these compounds have the molecular formula C_3H_6O and therefore the same molecular weight. Yet acetone and propylene oxide have distinctly different boiling points and chemical reactivity that, as a result, lend themselves to distinctly different practical applications. Their shared molecular formula simply gives us no basis for understanding the differences between them. We must, therefore, move to a consideration of their **structural formulas**.

Helpful Hint

Terms and concepts that are fundamentally important to your learning organic chemistry are set in bold blue type. You should learn them as they are introduced. These terms are also defined in the glossary.

Helpful Hint

Build handheld models of these compounds and compare their structures.

Acetone is used in some nail polish removers.

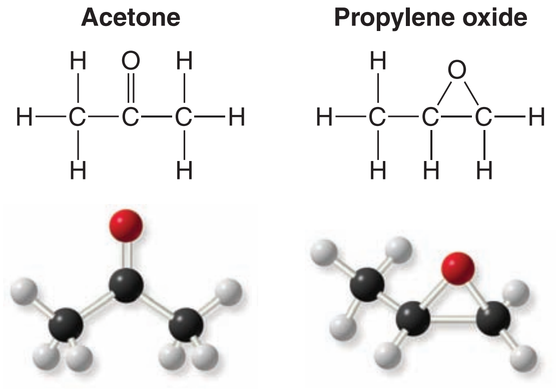


Figure 1.2 Ball-and-stick models and chemical formulas show the different structures of acetone and propylene oxide.

Propylene oxide alginates, made from propylene oxide and seaweed extracts, are used as food thickeners.

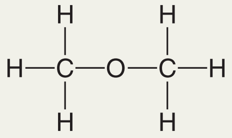
On examining the structures of acetone and propylene oxide several key aspects are clearly different (Fig. 1.2). Acetone contains a double bond between the oxygen atom and the central carbon atom. Propylene oxide does not contain a double bond, but has three atoms joined in a ring. The connectivity of the atoms is clearly different in acetone and propylene oxide. Their structures have the same molecular formula but a different constitution. We call such compounds constitutional isomers.*

- **Constitutional isomers** are different compounds that have the same molecular formula but differ in the sequence in which their atoms are bonded, that is, in their connectivity.
- Constitutional isomers usually have different physical properties (e.g., melting point, boiling point, and density) and different chemical properties (reactivity).

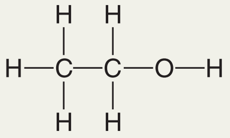
Solved Problem 1.1

There are two constitutional isomers with the formula C_2H_6O. Write structural formulas for these isomers.

STRATEGY AND ANSWER If we recall that carbon can form four covalent bonds, oxygen can form two, and hydrogen only one, we can arrive at the following constitutional isomers.



Dimethyl ether



Ethanol

Ethanol is the alcohol of alcoholic beverages.

It should be noted that these two isomers are clearly different in their physical properties. At room temperature and 1 atm pressure, dimethyl ether is a gas. Ethanol is a liquid.

1.3B The Tetrahedral Shape of Methane

In 1874, the structural formulas originated by Kekulé, Couper, and Butlerov were expanded into three dimensions by the independent work of J. H. van't Hoff and J. A. Le Bel. van't Hoff and Le Bel proposed that the four bonds of the carbon atom in methane, for example, are arranged in such a way that they would point toward the corners of a regular tetrahe-

---

*An older term for isomers of this type was **structural isomers**. The International Union of Pure and Applied Chemistry (IUPAC) now recommends that use of the term "structural" when applied to constitutional isomers be abandoned.

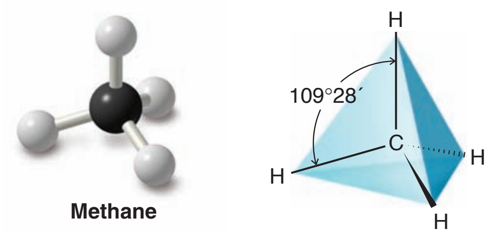


Figure 1.3 The tetrahedral structure of methane.

dron, the carbon atom being placed at its center (Fig. 1.3). The necessity for knowing the arrangement of the atoms in space, taken together with an understanding of the order in which they are connected, is central to an understanding of organic chemistry, and we shall have much more to say about this later, in Chapters 4 and 5.

1.4 *Chemical Bonds: The Octet Rule*

The first explanations of the nature of chemical bonds were advanced by G. N. Lewis (of the University of California, Berkeley) and W. Kössel (of the University of Munich) in 1916. Two major types of chemical bonds were proposed:

1. **Ionic** (or electrovalent) bonds are formed by the transfer of one or more electrons from one atom to another to create ions.
2. **Covalent** bonds result when atoms share electrons.

The central idea in their work on bonding is that atoms without the electronic configuration of a noble gas generally react to produce such a configuration because these configurations are known to be highly stable. For all of the noble gases except helium, this means achieving an octet of electrons in the valence shell.

- The tendency for an atom to achieve a configuration where its valence shell contains eight electrons is called the **octet rule**.

The concepts and explanations that arise from the original propositions of Lewis and Kössel are satisfactory for explanations of many of the problems we deal with in organic chemistry today. For this reason we shall review these two types of bonds in more modern terms.

1.4A Ionic Bonds

Atoms may gain or lose electrons and form charged particles called **ions**.

- An **ionic bond** is an attractive force between oppositely charged ions.

One source of such ions is a reaction between atoms of widely differing electronegativities (Table 1.2).

TABLE 1.2 Electronegativities of Some of the Elements

Increasing electronegativity →

		H 2.1				
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8						Br 2.8

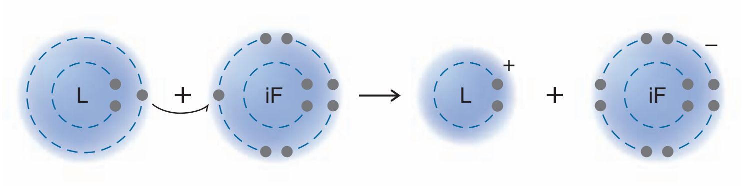
↑ Increasing electronegativity

Helpful Hint

We will use electronegativity frequently as a tool for understanding the properties and reactivity of organic molecules.

- **Electronegativity** ***is a measure of the ability of an atom to attract electrons.***
- Electronegativity increases as we go across a horizontal row of the periodic table from left to right and it increases as we go up a vertical column (Table 1.2).

An example of the formation of an ionic bond is the reaction of lithium and fluorine atoms:



Lithium, a typical metal, has a very low electronegativity; fluorine, a nonmetal, is the most electronegative element of all. The loss of an electron (a negatively charged species) by the lithium atom leaves a lithium cation (Li^+); the gain of an electron by the fluorine atom gives a fluoride anion (F^-).

- Ions form because atoms can achieve the electronic configuration of a noble gas by gaining or losing electrons.

The lithium cation with two electrons in its valence shell is like an atom of the noble gas helium, and the fluoride anion with eight electrons in its valence shell is like an atom of the noble gas neon. Moreover, crystalline lithium fluoride forms from the individual lithium and fluoride ions. In this process negative fluoride ions become surrounded by positive lithium ions, and positive lithium ions by negative fluoride ions. In this crystalline state, the ions have substantially lower energies than the atoms from which they have been formed. Lithium and fluorine are thus "stabilized" when they react to form crystalline lithium fluoride.

We represent the formula for lithium fluoride as LiF, because that is the simplest formula for this ionic compound.

Ionic substances, because of their strong internal electrostatic forces, are usually very high melting solids, often having melting points above 1000°C. In polar solvents, such as water, the ions are solvated (see Section 2.13D), and such solutions usually conduct an electric current.

- Ionic compounds, often called **salts**, form only when atoms of very different electronegativities transfer electrons to become ions.

1.4B Covalent Bonds and Lewis Structures

When two or more atoms of the same or similar electronegativities react, a complete transfer of electrons does not occur. In these instances the atoms achieve noble gas configurations by *sharing electrons*.

- **Covalent bonds** form by sharing of electrons between atoms of similar electronegativities to achieve the configuration of a noble gas.
- **Molecules** are composed of atoms joined exclusively or predominantly by covalent bonds.

Molecules may be represented by electron-dot formulas or, more conveniently, by bond formulas where each pair of electrons shared by two atoms is represented by a line. Some examples are shown here:

1. Hydrogen, being in group IA of the periodic table, has one valence electron. Two hydrogen atoms share electrons to form a hydrogen molecule, H_2.

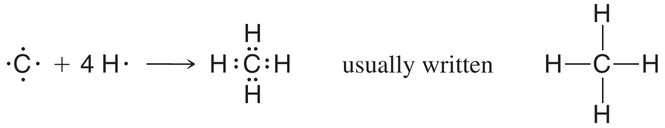
$$H_2 \qquad H\cdot + \cdot H \longrightarrow H{:}H \quad \text{usually written} \quad H{-}H$$

2. Because chlorine is in group VIIA, its atoms have seven valence electrons. Two chlorine atoms can share electrons (one electron from each) to form a molecule of Cl_2.

Cl_2

3. And a carbon atom (group IVA) with four valence electrons can share each of these electrons with four hydrogen atoms to form a molecule of methane, CH_4.

CH_4



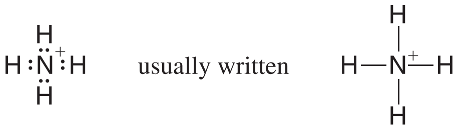
These formulas are often called **Lewis structures**; in writing them we show only the electrons of the valence shell.

4. Atoms can share *two or more pairs of electrons* to form **multiple covalent bonds**. For example, two nitrogen atoms possessing five valence electrons each (because nitrogen is in group VA) can share electrons to form a triple bond between them.

N_2

5. Ions, themselves, may contain covalent bonds. Consider, as an example, the ammonium ion.

$\overset{+}{N}H_4$



Review Problem 1.3

Consider the following compounds and decide whether the bond in them would be ionic or covalent.

(a) LiH **(b)** KCl **(c)** F_2 **(d)** PH_3

1.5 How to *Write Lewis Structures*

Several simple rules allow us to draw proper Lewis structures:

1. **Lewis structures show the connections between atoms in a molecule or ion using only the valence electrons of the atoms involved**. Valence electrons are those of an atom's outermost shell.

Helpful Hint

The ability to write proper **Lewis structures** is one of the most important tools for learning organic chemistry.

2. **For main group elements, the number of valence electrons a neutral atom brings to a Lewis structure is the same as its group number in the periodic table**. Carbon, for example, is in group IVA and has four valence electrons; the halogens (e.g., fluorine) are in group VIIA and each has seven valence electrons; hydrogen is in group IA and has one valence electron.
3. **If the structure we are drawing is a negative ion (an anion), we add one electron for each negative charge to the original count of valence electrons. If the structure is a positive ion (a cation), we subtract one electron for each positive charge**.
4. **In drawing Lewis structures we try to give each atom the electron configuration of a noble gas**. To do so, we draw structures where atoms share electrons to form covalent bonds or transfer electrons to form ions.
   a. Hydrogen forms one covalent bond by sharing its electron with an electron of another atom so that it can have two valence electrons, the same number as in the noble gas helium.

b. Carbon forms four covalent bonds by sharing its four valence electrons with four valence electrons from other atoms, so that it can have eight electrons (the same as the electron configuration of neon, satisfying the octet rule).

c. To achieve an octet of valence electrons, elements such as nitrogen, oxygen, and the halogens typically share only some of their valence electrons through covalent bonding, leaving others as unshared electron pairs.

The following problems illustrate this method.

Solved Problem 1.2

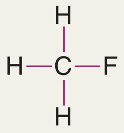
Write the Lewis structure of CH_3F.

STRATEGY AND ANSWER

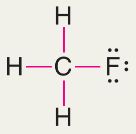
1. We find the total number of valence electrons of all the atoms:

$$\underset{\text{C}}{\underset{\uparrow}{4}} + \underset{3\,\text{H}}{\underset{\uparrow}{3(1)}} + \underset{\text{F}}{\underset{\uparrow}{7}} = 14$$

2. We use pairs of electrons to form bonds between all atoms that are bonded to each other. We represent these bonding pairs with lines. In our example this requires four pairs of electrons (8 of the 14 valence electrons).



3. We then add the remaining electrons in pairs so as to give each hydrogen 2 electrons (a duet) and every other atom 8 electrons (an octet). In our example, we assign the remaining 6 valence electrons to the fluorine atom in three nonbonding pairs.



Solved Problem 1.3

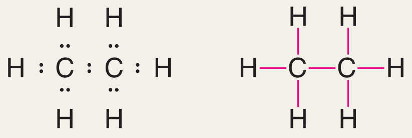
Write the Lewis structure for ethane (C_2H_6).

STRATEGY AND ANSWER

1. We find the total number of valence electrons of all the atoms.

$$\underset{2\,\text{C}}{\underset{\uparrow}{2(4)}} + \underset{6\,\text{H}}{\underset{\uparrow}{6(1)}} = 14$$

2. We use one pair of electrons to form a single bond between two carbon atoms, and six pairs of electrons to form single bonds from each carbon atom to three hydrogen atoms.



Solved Problem 1.4

Write a Lewis structure for methylamine (CH_5N).

STRATEGY AND ANSWER

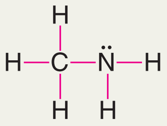
1. We find the total number of valence electrons for all the atoms.

$$\underset{\substack{\uparrow\\ \mathrm{C}}}{4}\quad \underset{\substack{\uparrow\\ \mathrm{N}}}{5}\quad \underset{\substack{\uparrow\\ 5\,\mathrm{H}}}{5(1)} = 14 = 7 \text{ pairs}$$

2. We use one electron pair to join the carbon and nitrogen.

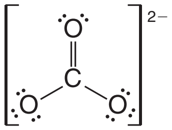
$$\mathrm{C{-}N}$$

3. We use three pairs to form single bonds between the carbon and three hydrogen atoms.
4. We use two pairs to form single bonds between the nitrogen atom and two hydrogen atoms.
5. This leaves one electron pair, which we use as a lone pair on the nitrogen atom.

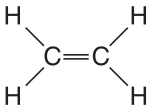


$$\mathrm{H{-}\overset{\displaystyle H}{\underset{\displaystyle H}{C}}{-}\underset{\displaystyle H}{\ddot{N}}{-}H}$$

If necessary, we use multiple bonds to satisfy the octet rule (i.e., give atoms the noble gas configuration). The carbonate ion (CO_3^{2-}) illustrates this:



The organic molecules ethene (C_2H_4) and ethyne (C_2H_2) have a double and triple bond, respectively:



and

$$\mathrm{H{-}C{\equiv}C{-}H}$$

1.6 *Exceptions to the Octet Rule*

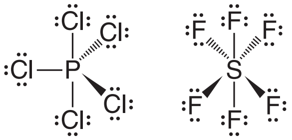
Atoms share electrons, not just to obtain the configuration of an inert gas, but because sharing electrons produces increased electron density between the positive nuclei. The resulting attractive forces of nuclei for electrons is the "glue" that holds the atoms together (cf. Section 1.11).

- Elements of the second period of the periodic table can have a maximum of four bonds (i.e., have eight electrons around them) because these elements have only one 2*s* and three 2*p* orbitals available for bonding.

Each orbital can contain two electrons, and a total of eight electrons fills these orbitals (Section 1.10A). The octet rule, therefore, only applies to these elements, and even here, as we shall see in compounds of beryllium and boron, fewer than eight electrons are possible.

- Elements of the third period and beyond have *d* orbitals that can be used for bonding.

These elements can accommodate more than eight electrons in their valence shells and therefore can form more than four covalent bonds. Examples are compounds such as PCl_5 and SF_6. Bonds written as ⋰ (dashed wedges) project behind the plane of the paper. Bonds written as ◢ (solid wedges) project in front of the paper.



Solved Problem 1.5

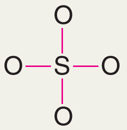
Write a Lewis structure for the sulfate ion (SO_4^{2-}). (*Note*: The sulfur atom is bonded to all four oxygen atoms.)

ANSWER

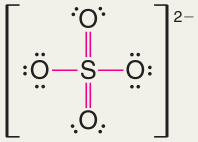
1. We find the total number of valence electrons including the extra 2 electrons needed to give the ion the double negative charge:

$$\underset{\underset{\text{S}}{\uparrow}}{6} + \underset{\underset{\text{4 O}}{\uparrow}}{4(6)} + \underset{\underset{2\,e^{-}}{\uparrow}}{2} = 32$$

2. We use four pairs of electrons to form bonds between the sulfur atom and the four oxygen atoms:



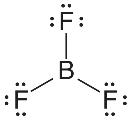
3. We add the remaining 24 electrons as unshared pairs on oxygen atoms and as double bonds between the sulfur atom and two oxygen atoms. This gives each oxygen 8 electrons and the sulfur atom 12:



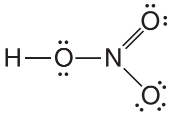
Review Problem 1.4

Write a Lewis structure for the phosphate ion (PO_4^{3-}).

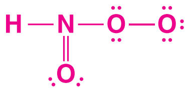
Some highly reactive molecules or ions have atoms with fewer than eight electrons in their outer shell. An example is boron trifluoride (BF_3). In a BF_3 molecule the central boron atom has only six electrons around it:



Finally, one point needs to be stressed: **Before we can write some Lewis structures, *we must know how the atoms are connected to each other***. Consider nitric acid, for example. Even though the formula for nitric acid is often written HNO_3, the hydrogen is actually connected to an oxygen, not to the nitrogen. The structure is $HONO_2$ and not HNO_3. Thus the correct Lewis structure is



and not



This knowledge comes ultimately from experiments. If you have forgotten the structures of some of the common inorganic molecules and ions (such as those listed in Review Problem 1.5), this may be a good time for a review of the relevant portions of your general chemistry text.

Helpful Hint

Check your progress by doing each Review Problem as you come to it in the text.

Solved Problem 1.6

Assume that the atoms are connected in the same way they are written in the formula, and write a Lewis structure for the toxic gas hydrogen cyanide (HCN).

STRATEGY AND ANSWER

1. We find the total number of valence electrons on all of the atoms:

$$\underset{\substack{\uparrow\\ \text{H}}}{1} + \underset{\substack{\uparrow\\ \text{C}}}{4} + \underset{\substack{\uparrow\\ \text{N}}}{5} = 10$$

2. We use one pair of electrons to form a single bond between the hydrogen atom and the carbon atom (see below), and we use three pairs to form a triple bond between the carbon atom and the nitrogen atom. This leaves two electrons. We use these as an unshared pair on the nitrogen atom. Now each atom has the electronic structure of a noble gas. The carbon atom has two electrons (like helium) and the carbon and nitrogen atoms each have eight electrons (like neon).

$$\text{H}{-}\text{C}{\equiv}\text{N}\colon$$

Review Problem 1.5

Write a Lewis structure for each of the following:

(a) HF
(b) F_2
(c) CH_3F
(d) HNO_2
(e) H_2SO_3
(f) BH_4^-
(g) H_3PO_4
(h) H_2CO_3

1.7 Formal Charges and How to Calculate Them

Many Lewis structures are incomplete until we decide whether any of their atoms have a **formal charge**. Calculating the formal charge on an atom in a Lewis structure is simply a bookkeeping method for its valence electrons.

Helpful Hint

Proper assignment of **formal charges** is another essential tool for learning organic chemistry.

- First, we examine each atom and, using the periodic table, we determine how many **valence electrons** it would have if it were an atom not bonded to any other atoms. **This is equal to the group number of the atom in the periodic table**. For hydrogen this number equals 1, for carbon it equals 4, for nitrogen it equals 5, and for oxygen it equals 6.

Next, we examine the atom in the Lewis structure and we assign the valence electrons in the following way:

- **We assign to each atom half of the electrons it is sharing with another atom and all of its unshared (lone) electron pairs**.

Then we do the following calculation for the atom:

Formal charge = number of valence electrons − 1/2 number of shared electrons –
number of unshared electrons

or

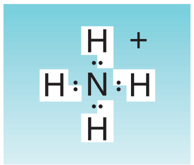
$$F = Z - (1/2)S - U$$

where F is the formal charge, Z is the group number of the element, S equals the number of shared electrons, and U is the number of unshared electrons.

- It is important to note, too, that **the arithmetic sum of all the formal charges in a molecule or ion will equal the overall charge on the molecule or ion**.

Let us consider several examples showing how this is done.

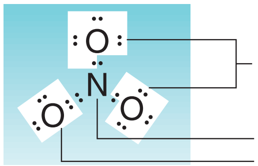
The Ammonium Ion (NH_4^+) As we see below, the ammonium ion has no unshared electron pairs. We divide all of the electrons in bonds equally between the atoms that share them. Thus, each hydrogen is assigned one electron. We subtract this from one (the number of valence electrons in a hydrogen atom) to give each hydrogen atom a formal charge of zero. The nitrogen atom is assigned four electrons (one from each bond). We subtract four from five (the number of valence electrons in a nitrogen atom) to give the nitrogen a formal charge of +1.



For hydrogen:	valence electrons of free atom	=	1
	subtract assigned electrons	=	– 1
	Formal charge on each hydrogen	=	0
For nitrogen:	valence electrons of free atom	=	5
	subtract assigned electrons	=	$-(\frac{1}{2})8$
	Formal charge on nitrogen	=	+1

Overall charge on ion = 4(0) + 1 = +1

The Nitrate Ion (NO_3^-) Let us next consider the nitrate ion (NO_3^-), an ion that has oxygen atoms with unshared electron pairs. Here we find that the nitrogen atom has a formal charge of +1, that two oxygen atoms have formal charges of −1, and that one oxygen has a formal charge equal to 0.



Formal charge = $6 - (\frac{1}{2})2 - 6 = -1$

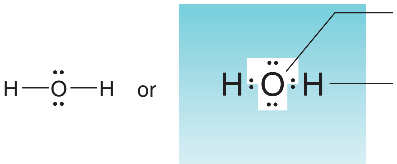
Formal charge = $5 - (\frac{1}{2})8 = +1$

Formal charge = $6 - (\frac{1}{2})4 - 4 = 0$

Charge on ion = 2(–1) + 1 + 0 = –1

Water and Ammonia The sum of the formal charges on each atom making up a molecule must be zero. Consider the following examples:

Water

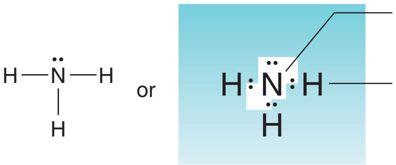


Formal charge = $6 - (\frac{1}{2})4 - 4 = 0$

Formal charge = $1 - (\frac{1}{2})2 = 0$

Charge on molecule = 0 + 2(0) = 0

Ammonia



Formal charge = $5 - (\frac{1}{2})6 - 2 = 0$

Formal charge = $1 - (\frac{1}{2})2 = 0$

Charge on molecule = 0 + 3(0) = 0

Review Problem 1.6

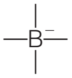
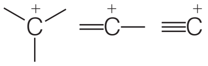
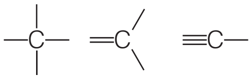
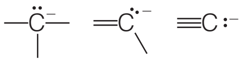
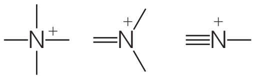
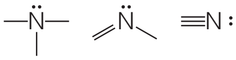
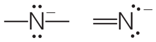
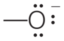
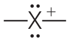
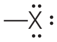
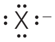
Write a Lewis structure for each of the following negative ions, and assign the formal negative charge to the correct atom:

(a) CH_3O^- **(c)** CN^- **(e)** HCO_3^-
(b) NH_2^- **(d)** HCO_2^- **(f)** HC_2^-

1.7A A Summary of Formal Charges

With this background, it should now be clear that each time an oxygen atom of the type $-\ddot{\underset{\cdot\cdot}{\mathrm{O}}}\colon$ appears in a molecule or ion, it will have a formal charge of − 1, and that each time an oxygen atom of the type $=\!\ddot{\underset{\cdot\cdot}{\mathrm{O}}}$ or $-\ddot{\underset{\cdot\cdot}{\mathrm{O}}}-$ appears, it will have a formal charge of 0. Similarly, $-\overset{|}{\underset{|}{\mathrm{N}}}-$ will be + 1, and $-\ddot{\underset{|}{\mathrm{N}}}-$ will be zero. These and other common structures are summarized in Table 1.3.

TABLE 1.3 A Summary of Formal Charges

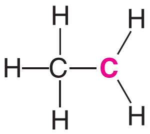
Group	Formal Charge of +1	Formal Charge of 0	Formal Charge of −1
IIIA			
IVA			
VA			
VIA			
VIIA		 (X = F, Cl, Br, or I)	

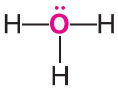
Helpful Hint

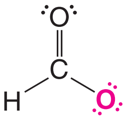
In later chapters, when you are evaluating how reactions proceed and what products form, you will find it essential to keep track of formal charges.

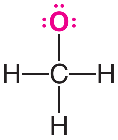
Review Problem 1.7

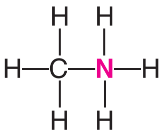
Assign the proper formal charge to the colored atom in each of the following structures:

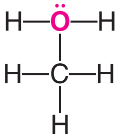
(a) 

(b) 

(c) 

(d) 

(e) 

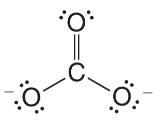
(f) 

(g) $CH_3-C\equiv N\colon$

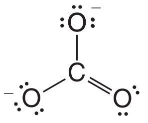
(h) $CH_3-N\equiv N\colon$

1.8 Resonance Theory

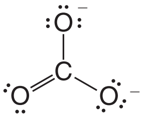
Many times more than one *equivalent* Lewis structure can be written for a molecule or ion. Consider, for example, the carbonate ion (CO_3^{2-}). We can write three *different* but *equivalent* structures, **1–3**:



1



2



3

Notice two important features of these structures. First, each atom has the noble gas configuration. Second, *and this is especially important*, we can convert one structure into any other by *changing only the positions of the electrons*. We do not need to change the relative positions of the atomic nuclei. For example, if we move the electron pairs in the manner indicated by the **curved arrows** in structure **1**, we change structure **1** into structure **2**:

1 **2**

Helpful Hint

Curved arrows (Section 3.5) show movement of electron pairs, *not atoms*. The *tail* of the arrow begins at the current position of the electron pair. The *head* of the arrow points to the location where the electron pair will be in the next structure. Curved-arrow notation is one of the most important tools that you will use to understand organic reactions.

In a similar way we can change structure **2** into structure **3**:

2 **3**

Structures **1–3**, although not identical on paper, *are equivalent*. None of them alone, however, fits important data about the carbonate ion.

X-Ray studies have shown that carbon–oxygen double bonds are shorter than single bonds. The same kind of study of the carbonate ion shows, however, that all of its carbon–oxygen bonds *are of equal length*. One is not shorter than the others as would be expected from representations **1**, **2**, and **3**. Clearly none of the three structures agrees with this evidence. In each structure, **1–3**, one carbon–oxygen bond is a double bond and the other two are single bonds. None of the structures, therefore, is correct. How, then, should we represent the carbonate ion?

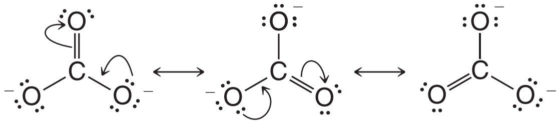
One way is through a theory called **resonance theory**. This theory states that whenever a molecule or ion can be represented by two or more Lewis structures *that differ only in the positions of the electrons*, two things will be true:

1. None of these structures, which we call **resonance structures** or **resonance contributors**, will be a realistic representation for the molecule or ion. None will be in complete accord with the physical or chemical properties of the substance.
2. The actual molecule or ion will be better represented by a *hybrid (average) of these structures*.

- ***Resonance structures, then, are not real structures for the actual molecule or ion; they exist only on paper***. As such, they can never be isolated. No single contributor adequately represents the molecule or ion. In resonance theory we view the carbonate ion, which is, of course, a real entity, as having a structure that is a **hybrid** of the three **hypothetical** resonance structures.

What would a hybrid of structures **1–3** be like? Look at the structures and look especially at a particular carbon–oxygen bond, say, the one at the top. This carbon–oxygen bond is a double bond in one structure (**1**) and a single bond in the other two (**2** and **3**). The actual carbon–oxygen bond, since it is a hybrid, must be something in between a double bond and a single bond. Because the carbon–oxygen bond is a single bond in two of the structures and a double bond in only one, it must be more like a single bond than a double bond. It must be like a one and one-third bond. We could call it a partial double bond. And, of course, what we have just said about any one carbon–oxygen bond will be equally true of the other two. Thus all of the carbon–oxygen bonds of the carbonate ion are partial double bonds, and *all are equivalent*. All of them *should be* the same length, and this is exactly what experiments tell us. The bonds are all 1.28 Å long, a distance which is intermediate between that of a carbon–oxygen single bond (1.43 Å) and that of a carbon–oxygen double bond (1.20 Å). One angstrom equals 1×10^{-10} meter.

- One other important point: By convention, when we draw resonance structures, we connect them by double-headed arrows (⟷) to indicate clearly that they are hypothetical, not real. For the carbonate ion we write them this way:



We should not let these arrows, or the word "resonance," mislead us into thinking that the carbonate ion fluctuates between one structure and another. These structures individually do not represent reality and exist only on paper; therefore, the carbonate ion cannot fluctuate among them because it is a hybrid of them.

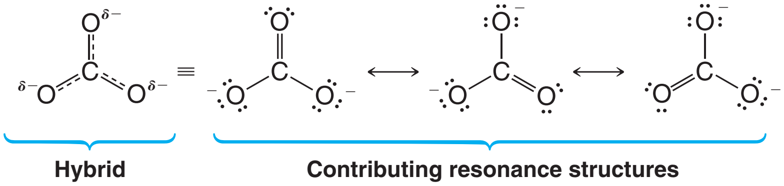
Helpful Hint

Each type of arrow in organic chemistry (e.g., ⌒, ⇋, and ⟷) has a specific meaning. It is important that you use each type of arrow only for the purpose for which it is defined.

- It is also important to distinguish between resonance and an **equilibrium**.

In an equilibrium between two, or more, species, it is quite correct to think of different structures and moving (or fluctuating) atoms, *but not in the case of resonance* (as in the carbonate ion). Here the atoms do not move, and the "structures" exist only on paper. **An equilibrium is indicated by ⇋ and resonance by ⟷.**

How can we write the structure of the carbonate ion in a way that will indicate its actual structure? We may do two things: we may write all of the resonance structures as we have just done and let the reader mentally fashion the hybrid, or we may write a non-Lewis structure that attempts to represent the hybrid. For the carbonate ion we might do the following:



The bonds in the structure on the left are indicated by a combination of a solid line and a dashed line. This is to indicate that the bonds are something in between a single bond and a double bond. As a rule, we use a solid line whenever a bond appears in all structures, and a dashed line when a bond exists in one or more but not all. We also place a $\delta-$ (read partial minus) beside each oxygen to indicate that something less than a full negative charge resides on each oxygen atom. (In this instance each oxygen atom has two-thirds of a full negative charge.)

Calculations from theory show the equal charge density at each oxygen in the carbonate anion. Figure 1.4 shows a calculated **electrostatic potential map** of the electron density in the carbonate ion. In an electrostatic potential map, regions of relatively more negative charge are red, while more positive regions (i.e., less negative regions) are indicated by colors trending toward blue. Equality of the bond lengths in the carbonate anion (partial double bonds as shown in the resonance hybrid above) is also evident in this model.

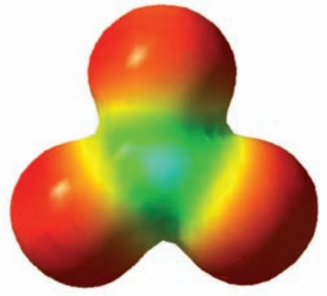
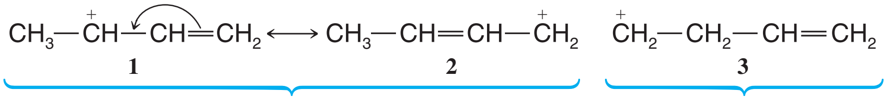


Figure 1.4 A calculated electrostatic potential map for the carbonate anion, showing the equal charge distribution at the three oxygen atoms. In electrostatic potential maps like this one, colors trending toward red mean increasing concentration of negative charge, while those trending toward blue mean less negative (or more positive) charge.

1.8A How to Write Resonance Structures

1. **Resonance structures exist only on paper**. Although they have no real existence of their own, **resonance structures** are useful because they allow us to describe molecules and ions for which a single Lewis structure is inadequate. We write two or more Lewis structures, calling them resonance structures or resonance contributors. We connect these structures by double-headed arrows (⟷), and we say that the real molecule or ion is a hybrid of all of them.
2. **We are only allowed to move electrons in writing resonance structures**. The positions of the nuclei of the atoms must remain the same in all of the structures.

Structure **3** is not a resonance structure of **1** or **2**, for example, because in order to form it we would have to move a hydrogen atom and this is not permitted:



$$CH_3-\overset{+}{C}H-CH=CH_2 \longleftrightarrow CH_3-CH=CH-\overset{+}{C}H_2 \qquad \overset{+}{C}H_2-CH_2-CH=CH_2$$

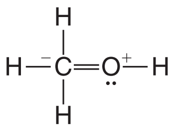
1 **2** **3**

These are resonance structures.

This is not a proper resonance structure of 1 or 2 because a hydrogen atom has been moved.

Generally speaking, when we move electrons, we move only those of multiple bonds (as in the example above) and those of nonbonding electron pairs.

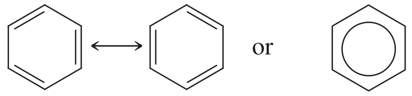
3. **All of the structures must be proper Lewis structures**. We should not write structures in which carbon has five bonds, for example:



This is not a proper resonance structure for methanol because carbon has five bonds. Elements of the first major row of the periodic table cannot have more than eight electrons in their valence shell.

4. **The energy of the resonance hybrid is lower than the energy of any contributing structure**. Resonance stabilizes a molecule or ion. This is especially true when the resonance structures are equivalent. Chemists call this stabilization *resonance stabilization*. **If the resonance structures are equivalent, then the resonance stabilization is large**.

In Chapter 14 we shall find that benzene is highly resonance stabilized because it is a hybrid of the two equivalent forms that follow:



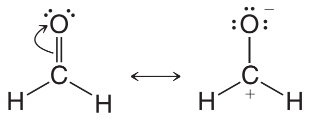
Resonance structures for benzene **Representation of hybrid**

5. **The more stable a structure is (when taken by itself), the greater is its contribution to the hybrid**.

How do we decide whether one resonance structure is more stable than another? The following rules will help us:

1. **The more covalent bonds a structure has, the more stable it is**. Consider the resonance structures for formaldehyde below. (Formaldehyde is a chemical used to preserve biological specimens.) Structure **A** has more covalent bonds, and therefore makes a larger contribution to the hybrid. In other words, the hybrid is more like structure **A** than structure **B**.

Four covalent bonds



Three covalent bonds

A **B**

More stable Less stable

Resonance structures for formaldehyde

These structures also illustrate two other considerations:

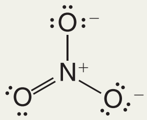
2. **Charge separation decreases stability**. It takes energy to separate opposite charges, and therefore a structure with separated charges is less stable. Structure **B** for

formaldehyde has separated plus and minus charges; therefore, on this basis, too, it is the less stable contributor and makes a smaller contribution to the hybrid.

3. **Structures in which all the atoms have a complete valence shell of electrons (i.e., the noble gas structure) are more stable**. Look again at structure **B**. The carbon atom has only six electrons around it, whereas in **A** it has eight. On this basis we can conclude that **A** is more stable and makes a large contribution.

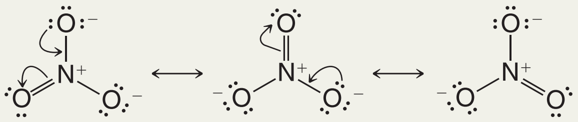
Solved Problem 1.7

The following is one way of writing the structure of the nitrate ion:

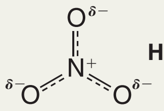


However, considerable physical evidence indicates that all three nitrogen–oxygen bonds are equivalent and that they have the same length, a bond distance between that expected for a nitrogen–oxygen single bond and a nitrogen–oxygen double bond. Explain this in terms of resonance theory.

STRATEGY AND ANSWER We recognize that if we move the electron pairs in the following way, we can write three *different* but *equivalent* structures for the nitrate ion:



Since these structures differ from one another *only in the positions of their electrons*, they are *resonance structures* or *resonance contributors*. As such, no single structure taken alone will adequately represent the nitrate ion. The actual molecule will be best represented by a *hybrid of these three structures*. We might write this hybrid in the following way to indicate that all of the bonds are equivalent and that they are more than single bonds and less than double bonds. We also indicate that each oxygen atom bears an equal partial negative charge. This charge distribution corresponds to what we find experimentally.



Hybrid structure for the nitrate ion

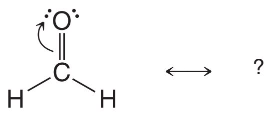
Review Problem 1.8

(a) Write two resonance structures for the formate ion HCO_2^-. (*Note*: The hydrogen and oxygen atoms are bonded to the carbon.) **(b)** Explain what these structures predict for the carbon–oxygen bond lengths of the formate ion, and **(c),** for the electrical charge on the oxygen atoms.

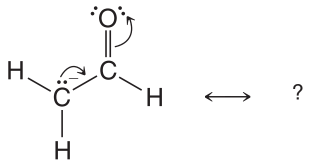
Review Problem 1.9

Write the resonance structure that would result from moving the electrons as the curved arrows indicate. Be sure to include formal charges if needed.

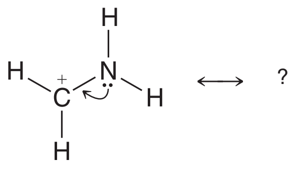
(a)



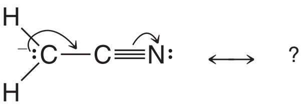
(b)



(c)



(d)



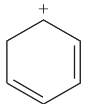
Review Problem 1.10

Write the contributing resonance structures and resonance hybrid for each of the following:

(a) $CH_3CH{=}CH{-}CH{=}\overset{+}{\underset{\cdot\cdot}{O}}H$

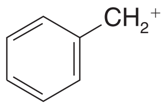
(b) $CH_2{=}CH{-}\underset{+}{C}H{-}CH{=}CH_2$

(c)

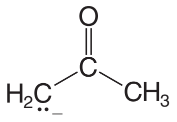


(d) $CH_2{=}CH{-}Br$

(e)



(f)



(g) $CH_3{-}S{-}CH_2^+$

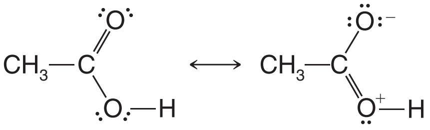
(h) $CH_3{-}NO_2$

Review Problem 1.11

From each set of resonance structures that follow, designate the one that would contribute most to the hybrid and explain your choice:

(a) $\overset{+}{C}H_2{-}\ddot{N}(CH_3)_2 \longleftrightarrow CH_2{=}\overset{+}{N}(CH_3)_2$

(b)



(c) $:NH_2{-}C{\equiv}N: \longleftrightarrow \overset{+}{N}H_2{=}C{=}\overset{\cdot\cdot}{\underset{\cdot\cdot}{N}}^{-}$

1.9 Quantum Mechanics and Atomic Structure

A theory of atomic and molecular structure was advanced independently and almost simultaneously by three people in 1926: Erwin Schrödinger, Werner Heisenberg, and Paul Dirac. This theory, called **wave mechanics** by Schrödinger and **quantum mechanics** by Heisenberg, has become the basis from which we derive our modern understanding of bonding in molecules. At the heart of quantum mechanics are equations called wave functions (denoted by the Greek letter psi, ψ).

- Each **wave function (ψ)** corresponds to a different *energy state* for an electron.
- Each *energy state* is a sublevel where one or two electrons can reside.
- The **energy** associated with the state of an electron can be calculated from the wave function.
- The **relative probability** of finding an electron in a given region of space can be calculated from the wave function (Section 1.10).
- The solution to a wave function can be positive, negative, or zero (Fig. 1.5).

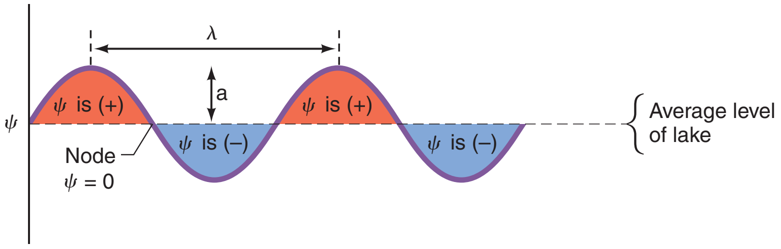
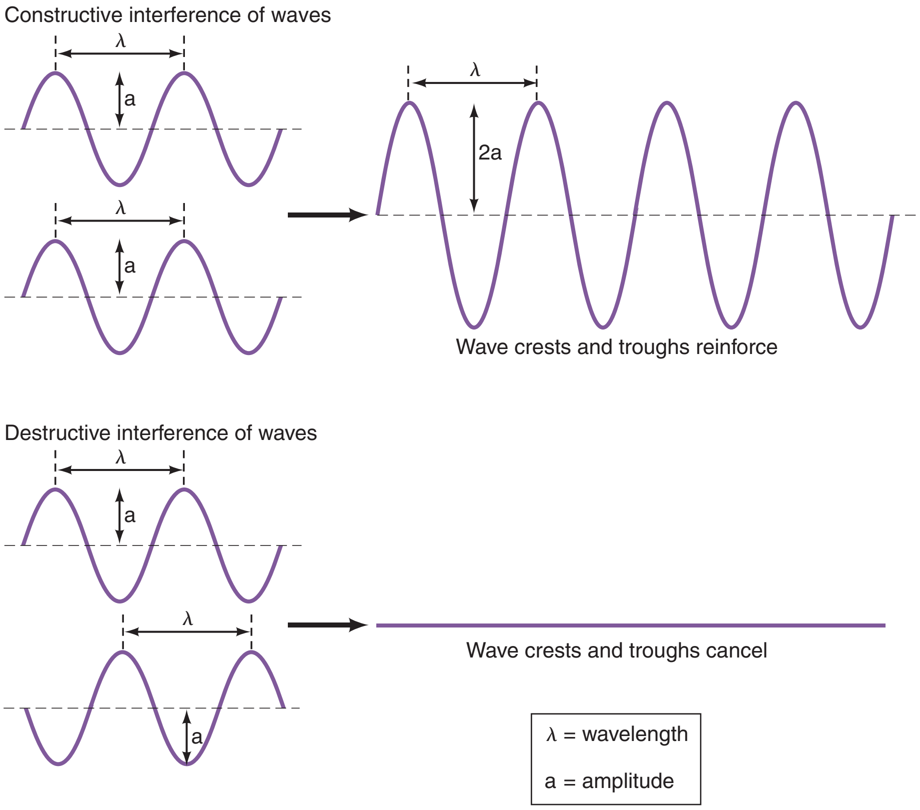


Figure 1.5 A wave moving across a lake is viewed along a slice through the lake. For this wave the wave function, ψ, is plus (+) in crests and minus (−) in troughs. At the average level of the lake it is zero; these places are called nodes. The magnitude of the crests and troughs is the amplitude (a) of the wave. The distance from the crest of one wave to the crest of the next is the wavelength (λ, or lambda).

- The **phase sign** of a wave equation indicates whether the solution is positive or negative when calculated for a given point in space relative to the nucleus.

Wave functions, whether they are for sound waves, lake waves, or the energy of an electron, have the possibility of constructive interference and destructive interference.

- **Constructive interference** occurs when wave functions with the same phase sign interact. There is a *reinforcing effect* and the amplitude of the wave function increases.
- **Destructive interference** occurs when wave functions with opposite phase signs interact. There is a *subtractive effect* and the amplitude of the wave function goes to zero or changes sign.



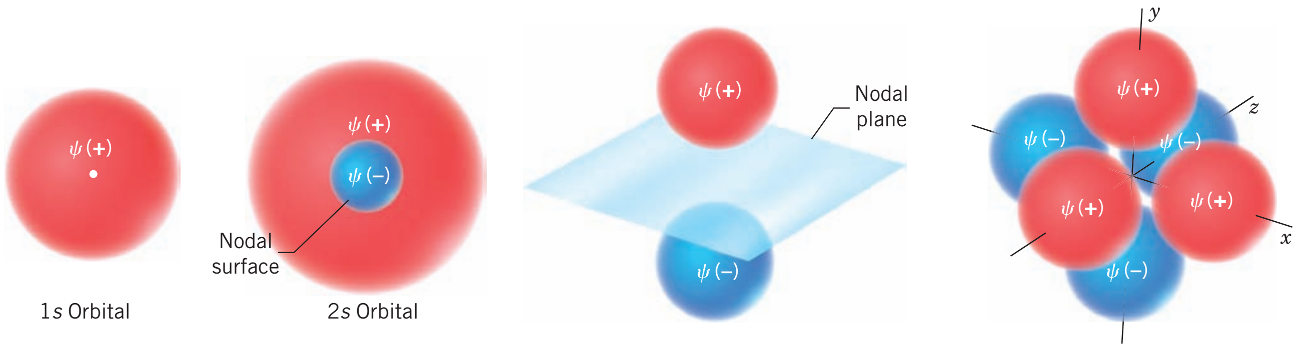
Experiments have shown that electrons have properties of waves and particles, which was an idea first put forth by Louis de Broglie in 1923. Our discussion focuses on the wavelike properties of electrons, however.

1.10 Atomic Orbitals and Electron Configuration

A physical interpretation related to the electron wave function was put forth by Max Born in 1926, as follows.

- The square of a wave function (ψ^2) for a particular x,y,z location expresses the probability of finding an electron at that location in space.

If the value of ψ^2 is large in a unit volume of space, the probability of finding an electron in that volume is high—we say that the **electron probability density** is large. Conversely, if ψ^2 for some other volume of space is small, the probability of finding an electron there



1*s* Orbital

2*s* Orbital

A pure (unhybridized) 2*p* orbital has two lobes.

There are three 2*p* orbitals, each with a (+) and (–) lobe, aligned symmetrically along the *x*, *y*, and *z* axes.

Figure 1.6 The shapes of some *s* and *p* orbitals. Pure, unhybridized *p* orbitals are almost-touching spheres. The *p* orbitals in hybridized atoms are lobe-shaped (Section 1.13).

is low.* This leads to the general definition of an orbital and, by extension, to the familiar shapes of atomic orbitals.

- An **orbital** is a region of space where the probability of finding an electron is high.
- **Atomic orbitals** are plots of ψ^2 in three dimensions. These plots generate the familiar *s*,*p*, and *d* orbital shapes.

The volumes that we show are those that would contain the electron 90–95% of the time. There is a finite, but very small, probability of finding an electron at greater distance from the nucleus than shown in the plots.

The shapes of *s* and *p* orbitals are shown in Fig. 1.6.

All ***s* orbitals** are spheres. A 1*s* orbital is a simple sphere. A 2*s* orbital is a sphere with an inner nodal surface ($\psi^2 = 0$). The inner portion of the 2*s* orbital, ψ_{2s}, has a negative phase sign.

The shape of a ***p* orbital** is like that of almost-touching spheres or lobes. The phase sign of a 2*p* wave function, ψ_{2p}, is positive in one lobe and negative in the other. A nodal plane separates the two lobes of a *p* orbital, and the three *p* orbitals of a given energy level are arranged in space along the *x, y*, and *z* axes in a Cartesian coordinate system.

- The + and − signs of wave functions do not imply positive or negative charge or greater or lesser probability of finding an electron.
- ψ^2 (the probability of finding an electron) is always positive, because squaring either a positive or negative solution to ψ leads to a positive value.

Thus, the probability of finding an electron in either lobe of a *p* orbital is the same. We shall see the significance of the + and − signs later when we see how atomic orbitals combine to form molecular orbitals.

1.10A Electron Configurations

The relative energies of atomic orbitals in the first and second principal shells are as follows:

- Electrons in 1*s* orbitals have the lowest energy because they are closest to the positive nucleus.
- Electrons in 2*s* orbitals are next lowest in energy.
- Electrons of the three 2*p* orbitals have equal but higher energy than the 2*s* orbital.

---

*Integration of ψ^2 over all space must equal 1; that is, the probability of finding an electron somewhere in all of space is 100%.

- Orbitals of equal energy (such as the three 2*p* orbitals) are called **degenerate orbitals**.

We can use these relative energies to arrive at the electron configuration of any atom in the first two rows of the periodic table. We need follow only a few simple rules.

1. **Aufbau principle**: Orbitals are filled so that those of lowest energy are filled first. (*Aufbau* is German for "building up.")
2. **Pauli exclusion principle**: A maximum of two electrons may be placed in each orbital *but only when the spins of the electrons are paired*. An electron spins about its own axis. For reasons that we cannot develop here, an electron is permitted only one or the other of just two possible spin orientations. We usually show these orientations by arrows, either ↑ or ↓. Thus two spin-paired electrons would be designated ↓↑. Unpaired electrons, which are not permitted in the same orbital, are designated ↑↑ (or ↓↓).
3. **Hund's rule**: When we come to orbitals of equal energy (degenerate orbitals) such as the three *p* orbitals, we add one electron to each *with their spins unpaired* until each of the degenerate orbitals contains one electron. (This allows the electrons, which repel each other, to be farther apart.) Then we begin adding a second electron to each degenerate orbital so that the spins are paired.

If we apply these rules to some of the second-row elements of the periodic table, we get the results shown in Fig. 1.7.

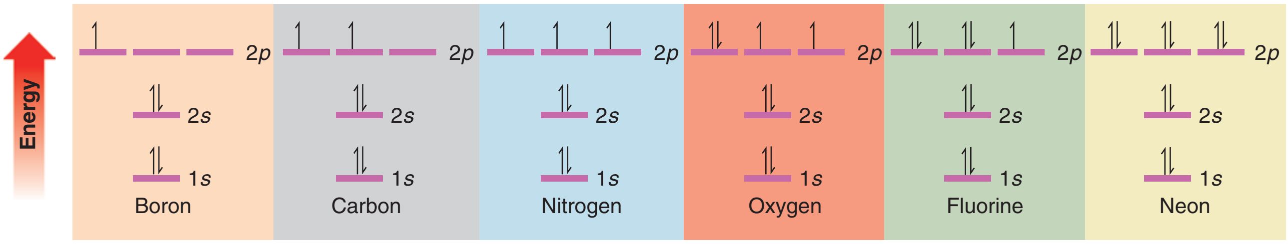


Figure 1.7 The ground state electron configurations of some second-row elements.

1.11 *Molecular Orbitals*

Atomic orbitals provide a means for understanding how atoms form covalent bonds. Let us consider a very simple case—formation of a bond between two hydrogen atoms to form a hydrogen molecule (Fig. 1.8).

When two hydrogen atoms are relatively far apart their total energy is simply that of two isolated hydrogen atoms (**I**). Formation of a covalent bond reduces the overall energy

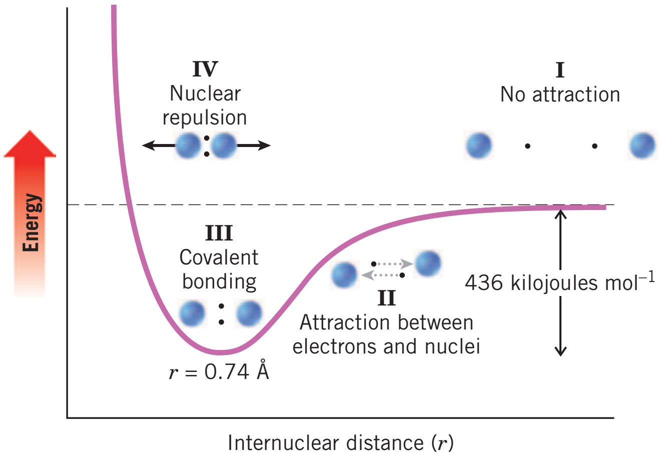


Figure 1.8 The potential energy of the hydrogen molecule as a function of internuclear distance.

of the system, however. As the two hydrogen atoms move closer together (**II**), each nucleus increasingly attracts the other's electron. This attraction more than compensates for the repulsive force between the two nuclei (or the two electrons). The result is a covalent bond (**III**), such that the internuclear distance is an ideal balance that allows the two electrons to be shared between both atoms while at the same time avoiding repulsive interactions between their nuclei. This ideal internuclear distance between hydrogen atoms is 0.74 Å, and we call this the **bond length** in a hydrogen molecule. If the nuclei are moved closer together (**IV**) the repulsion of the two positively charged nuclei predominates, and the energy of the system rises.

Notice that each H· has a shaded area around it, indicating that its precise position is uncertain. Electrons are constantly moving.

- According to the **Heisenberg uncertainty principle**, we cannot simultaneously know the position and momentum of an electron.

These shaded areas in our diagram represent orbitals, and they result from applying the principles of quantum mechanics. Plotting the square of the wave function (ψ^2) gives us a three-dimensional region called an orbital where finding an electron is highly probable.

- An **atomic orbital** represents the region of space where one or two electrons of an isolated atom are likely to be found.

In the case of our hydrogen model above, the shaded spheres represent the 1*s* orbital of each hydrogen atom. As the two hydrogen atoms approach each other their 1*s* orbitals begin to overlap until their atomic orbitals combine to form molecular orbitals.

- A **molecular orbital (MO)** represents the region of space where one or two electrons of a molecule are likely to be found.
- An orbital (atomic or molecular) can contain a maximum of two spin-paired electrons (Pauli exclusion principle).
- When atomic orbitals combine to form molecular orbitals, **the number of molecular orbitals that result always equals the number of atomic orbitals that combine**.

Thus, in the formation of a hydrogen molecule the two ψ_{1s} atomic orbitals combine to produce two molecular orbitals. Two orbitals result because the mathematical properties of wave functions permit them to be combined by either addition or subtraction. That is, they can combine either in or out of phase.

- A **bonding molecular orbital** (ψ_{molec}) results when two orbitals of the same phase overlap (Fig. 1.9).
- An **antibonding molecular orbital** (ψ^*_{molec}) results when two orbitals of opposite phase overlap (Fig. 1.10).

The bonding molecular orbital of a hydrogen molecule in its lowest energy (ground) state contains both electrons from the individual hydrogen atoms. The value of ψ (and therefore also ψ^2) is large between the nuclei, precisely as expected since the electrons are shared by both nuclei to form the covalent bond.

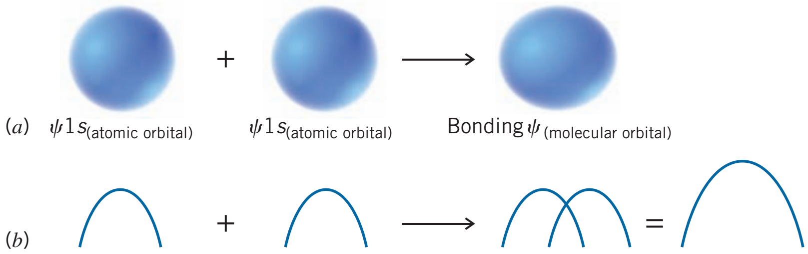


Figure 1.9 *(a)* The overlapping of two hydrogen 1*s* atomic orbitals with the same phase sign (indicated by their identical color) to form a bonding molecular orbital. *(b)* The analogous overlapping of two waves with the same phase, resulting in constructive interference and enhanced amplitude.

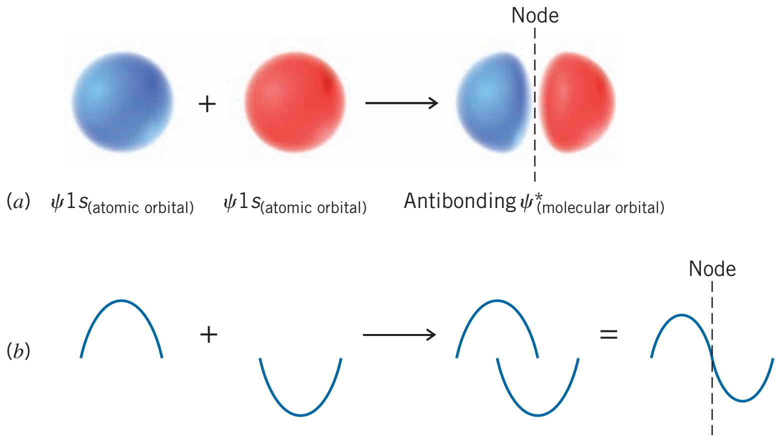


Figure 1.10 *(a)* The overlapping of two hydrogen 1*s* atomic orbitals with opposite phase signs (indicated by their different colors) to form an antibonding molecular orbital. *(b)* The analogous overlapping of two waves with the opposite sign, resulting in destructive interference and decreased amplitude. A node exists where complete cancellation by opposite phases makes the value of the combined wave function zero.

The antibonding molecular orbital contains no electrons in the ground state of a hydrogen molecule. Furthermore, the value of ψ (and therefore also ψ^2) goes to zero between the nuclei, creating a node ($\psi = 0$). The antibonding orbital does not provide for electron density between the atoms, and thus it is not involved in bonding.

What we have just described has its counterpart in a mathematical treatment called the **LCAO (linear combination of atomic orbitals)** method. In the LCAO treatment, wave functions for the atomic orbitals are combined in a linear fashion (by addition or subtraction) in order to obtain new wave functions for the molecular orbitals.

Molecular orbitals, like atomic orbitals, correspond to particular energy states for an electron. Calculations show that the relative energy of an electron in the bonding molecular orbital of the hydrogen molecule is substantially less than its energy in a ψ_{1s} atomic orbital. These calculations also show that the energy of an electron in the antibonding molecular orbital is substantially greater than its energy in a ψ_{1s} atomic orbital.

An energy diagram for the molecular orbitals of the hydrogen molecule is shown in Fig. 1.11. Notice that electrons are placed in molecular orbitals in the same way that they are in atomic orbitals. Two electrons (with their spins opposed) occupy the bonding molecular orbital, where their total energy is less than in the separate atomic orbitals. This is, as we have said, the *lowest electronic state* or *ground state* of the hydrogen molecule. An electron may occupy the antibonding molecular orbital in what is called an *excited state* for the molecule. This state forms when the molecule in the ground state (Fig. 1.11) absorbs a photon of light having the proper energy (ΔE).

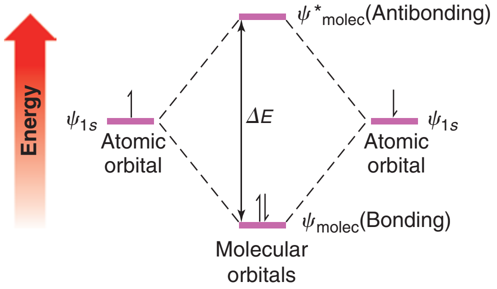


Figure 1.11 Energy diagram for the hydrogen molecule. Combination of two atomic orbitals, ψ_{1s}, gives two molecular orbitals, ψ_{molec} and ψ^*_{molec}. The energy of ψ_{molec} is lower than that of the separate atomic orbitals, and in the lowest electronic energy state of molecular hydrogen the bonding MO contains both electrons.

1.12 *The Structure of Methane and Ethane: sp^3 Hybridization*

The *s* and *p* orbitals used in the quantum mechanical description of the carbon atom, given in Section 1.10, were based on calculations for hydrogen atoms. These simple *s* and *p* orbitals do not, when taken alone, provide a satisfactory model for the *tetravalent–tetrahedral* carbon of methane (CH_4, see Review Problem 1.12). However, a satisfactory model of methane's structure that is based on quantum mechanics *can* be obtained through an approach called **orbital hybridization**. Orbital hybridization, in its simplest terms, is nothing more than a mathematical approach that involves the combining of individual wave functions for *s* and *p* orbitals to obtain wave functions for new orbitals. The new orbitals have, *in varying proportions*, the properties of the original orbitals taken separately. These new orbitals are called **hybrid atomic orbitals**.

According to quantum mechanics, the electronic configuration of a carbon atom in its lowest energy state—called the **ground state**—is that given here:

$$\text{C}\quad \underset{1s}{\underline{\upharpoonleft\downharpoonright}}\;\; \underset{2s}{\underline{\upharpoonleft\downharpoonright}}\;\; \underset{2p_x}{\underline{\upharpoonleft}}\;\; \underset{2p_y}{\underline{\upharpoonleft}}\;\; \underset{2p_z}{\underline{\quad}}$$

Ground state of a carbon atom

The valence electrons of a carbon atom (those used in bonding) are those of the *outer level*, that is, the 2*s* and 2*p* electrons.

1.12A The Structure of Methane

Hybrid atomic orbitals that account for the structure of methane can be derived from carbon's second-shell *s* and *p* orbitals as follows (Fig. 1.12):

- Wave functions for the 2*s*, $2p_x$, $2p_y$, and $2p_z$ orbitals of ground state carbon are mixed to form four new and equivalent $2sp^3$ hybrid orbitals.
- The designation sp^3 signifies that the hybrid orbital has one part *s* orbital character and three parts *p* orbital character.
- The mathematical result is that the four $2sp^3$ orbitals are oriented at angles of 109.5° with respect to each other. This is precisely the orientation of the four hydrogen atoms of methane. Each H—C—H bond angle is 109.5°.

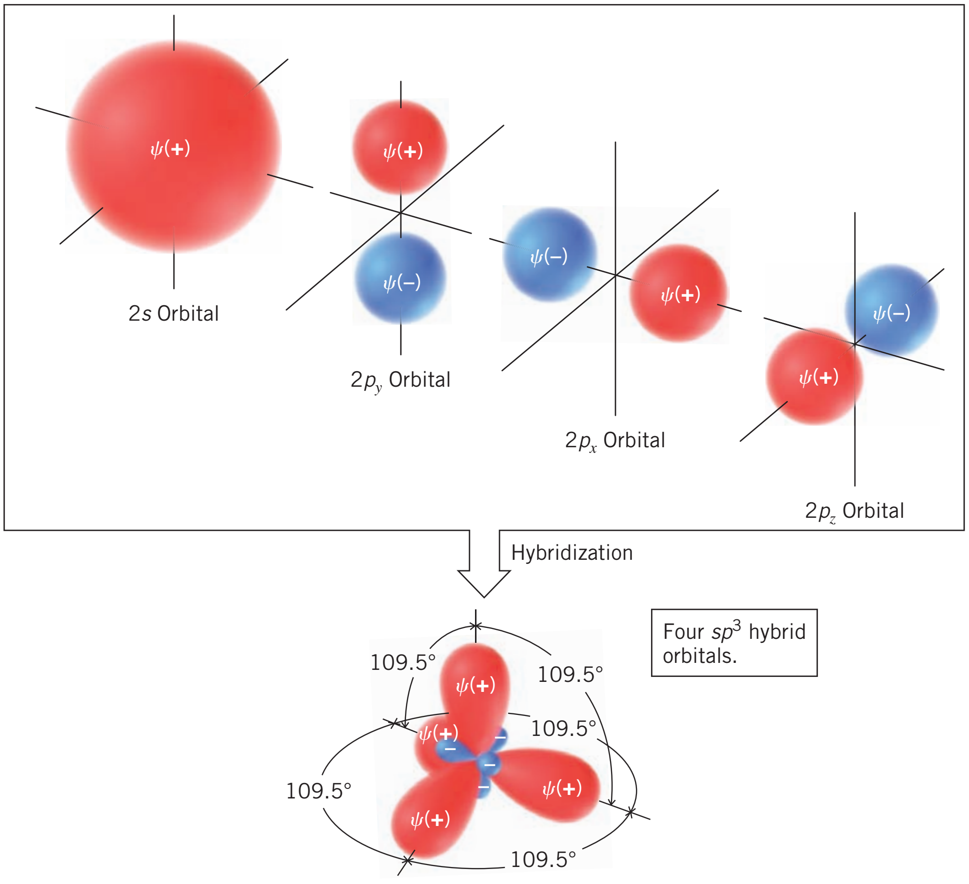
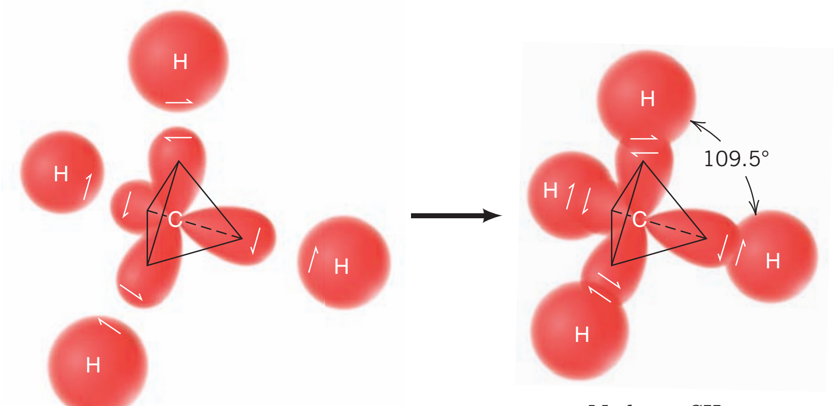


Figure 1.12 Hybridization of pure atomic orbitals of a carbon atom to produce sp^3 hybrid orbitals.



Methane, CH_4

Figure 1.13 The hypothetical formation of methane from an sp^3-hybridized carbon atom and four hydrogen atoms. In orbital hybridization we combine orbitals, *not* electrons. The electrons can then be placed in the hybrid orbitals as necessary for bond formation, but always in accordance with the Pauli principle of no more than two electrons (with opposite spin) in each orbital. In this illustration we have placed one electron in each of the hybrid carbon orbitals. In addition, we have shown only the bonding molecular orbital of each C—H bond because these are the orbitals that contain the electrons in the lowest energy state of the molecule.

If, in our imagination, we visualize the hypothetical formation of methane from an sp^3-hybridized carbon atom and four hydrogen atoms, the process might be like that shown in Fig. 1.13. For simplicity we show only the formation of the *bonding molecular orbital* for each carbon–hydrogen bond. We see that an sp^3-hybridized carbon gives a *tetrahedral structure for methane, and one with four equivalent C—H* bonds.

Review Problem 1.12

(a) Consider a carbon atom in its ground state. Would such an atom offer a satisfactory model for the carbon of methane? If not, why not? (*Hint*: Consider whether a ground state carbon atom could be tetravalent, and consider the bond angles that would result if it were to combine with hydrogen atoms.)

(b) Consider a carbon atom in the excited state:

$$\begin{array}{cccccc} \text{C} & \underline{\uparrow\downarrow} & \underline{\uparrow} & \underline{\uparrow} & \underline{\uparrow} & \underline{\uparrow} \\ & 1s & 2s & 2p_x & 2p_y & 2p_z \end{array}$$

Excited state of a carbon atom

Would such an atom offer a satisfactory model for the carbon of methane? If not, why not?

In addition to accounting properly for the shape of methane, the orbital hybridization model also explains the very strong bonds that are formed between carbon and hydrogen. To see how this is so, consider the shape of an individual sp^3 orbital shown in Fig. 1.14. Because an sp^3 orbital has the character of a *p* orbital, the positive lobe of an sp^3 orbital is large and extends relatively far from the carbon nucleus.

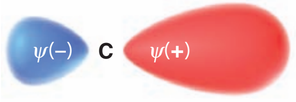


Figure 1.14 The shape of an sp^3 orbital.

It is the positive lobe of an sp^3 orbital that overlaps with the positive 1*s* orbital of hydrogen to form the bonding molecular orbital of a carbon–hydrogen bond (Fig. 1.15). Because the positive lobe of the sp^3 orbital is large and is extended into space, the overlap between it and the 1*s* orbital of hydrogen is also large, and the resulting carbon–hydrogen bond is quite strong.

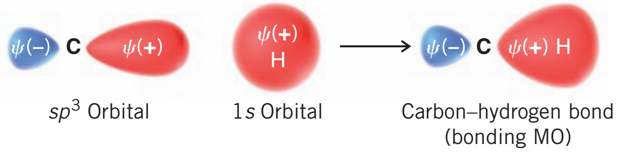


Figure 1.15 Formation of a C—H bond.

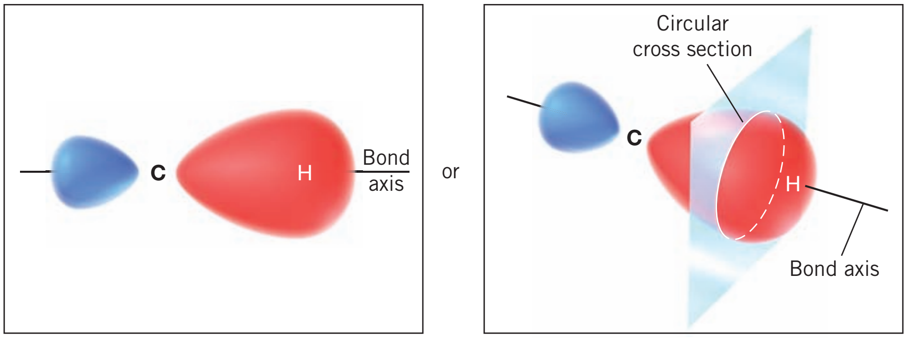


Figure 1.16 A σ (sigma) bond.

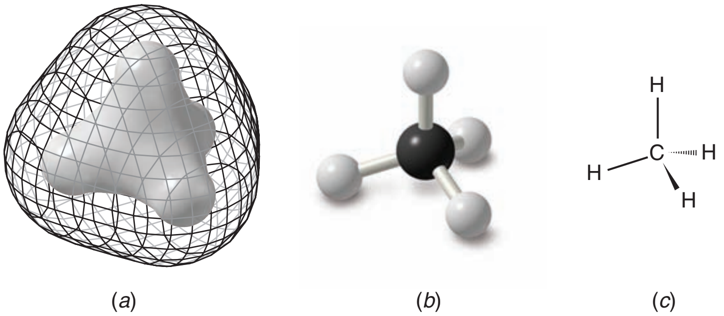
The bond formed from the overlap of an sp^3 orbital and a 1*s* orbital is an example of a **sigma (σ) bond** (Fig. 1.16).

- A **sigma (σ) bond** has a circularly symmetrical orbital cross section when viewed along the bond between two atoms.
- All purely **single bonds** are sigma bonds.

From this point on we shall often show only the bonding molecular orbitals because they are the ones that contain the electrons when the molecule is in its lowest energy state. Consideration of antibonding orbitals is important when a molecule absorbs light and in explaining certain reactions. We shall point out these instances later.

In Fig. 1.17 we show a calculated structure for methane where the tetrahedral geometry derived from orbital hybridization is clearly apparent.

Figure 1.17 *(a)* In this structure of methane, based on quantum mechanical calculations, the inner solid surface represents a region of high electron density. High electron density is found in each bonding region. The outer mesh surface represents approximately the furthest extent of overall electron density for the molecule. *(b)* This ball-and-stick model of methane is like the kind you might build with a molecular model kit. *(c)* This structure is how you would draw methane. Ordinary lines are used to show the two bonds that are in the plane of the paper, a solid wedge is used to show the bond that is in front of the paper, and a dashed wedge is used to show the bond that is behind the plane of the paper.



1.12B The Structure of Ethane

The bond angles at the carbon atoms of ethane, and of all alkanes, are also tetrahedral like those in methane. A satisfactory model for ethane can be provided by sp^3-hybridized carbon atoms. Figure 1.18 shows how we might imagine the bonding molecular orbitals of an ethane molecule being constructed from two sp^3-hybridized carbon atoms and six hydrogen atoms.

The carbon–carbon bond of ethane is a *sigma bond* with cylindrical symmetry, formed by two overlapping sp^3 orbitals. (The carbon–hydrogen bonds are also sigma bonds. They are formed from overlapping carbon sp^3 orbitals and hydrogen *s* orbitals.)

- Rotation of groups joined by a single bond does not usually require a large amount of energy.

Consequently, groups joined by single bonds rotate relatively freely with respect to one another. (We discuss this point further in Section 4.8.) In Fig. 1.19 we show a calculated structure for ethane in which the tetrahedral geometry derived from orbital hybridization is clearly apparent.

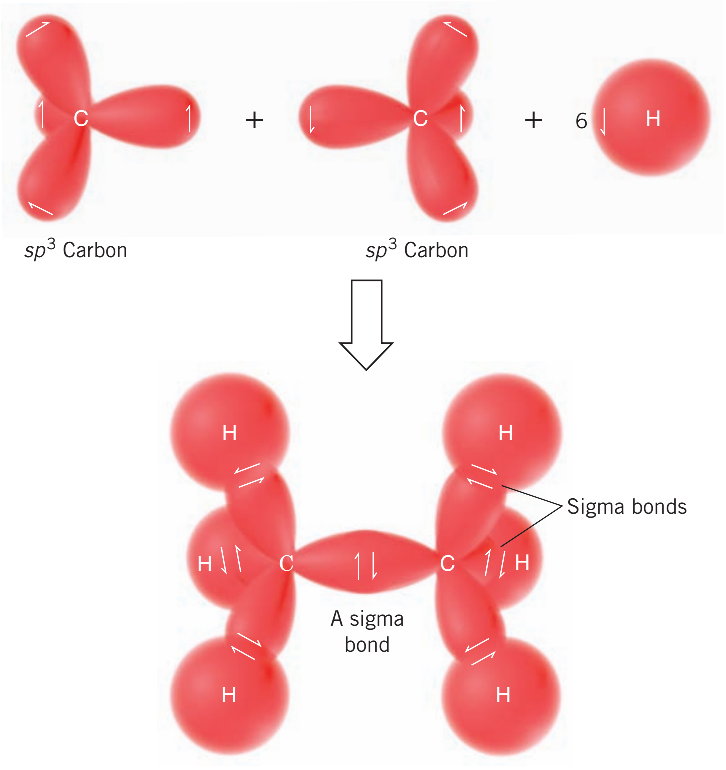


Figure 1.18 The hypothetical formation of the bonding molecular orbitals of ethane from two sp^3-hybridized carbon atoms and six hydrogen atoms. All of the bonds are sigma bonds. (Antibonding sigma molecular orbitals—called σ^* orbitals—are formed in each instance as well, but for simplicity these are not shown.)

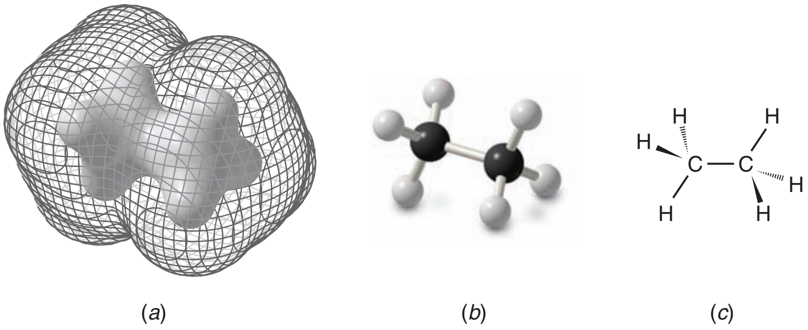
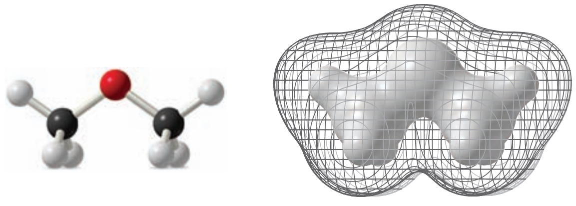


Figure 1.19 *(a)* In this structure of ethane, based on quantum mechanical calculations, the inner solid surface represents a region of high electron density. High electron density is found in each bonding region. The outer mesh surface represents approximately the furthest extent of overall electron density for the molecule. *(b)* A ball-and-stick model of ethane, like the kind you might build with a molecular model kit. *(c)* A structural formula for ethane as you would draw it using lines, wedges, and dashed wedges to show in three dimensions its tetrahedral geometry at each carbon.

THE CHEMISTRY OF . . .

Calculated Molecular Models: Electron Density Surfaces

We make frequent use in this book of molecular models derived from quantum mechanical calculations. These models will help us visualize the shapes of molecules as well as understand their properties and reactivity. A useful type of model is one that shows a calculated three-dimensional surface at which a chosen value of electron density is the same all around a molecule, called an **electron density surface**. If we make a plot where the value chosen is for low electron density, the result is a van der Waals surface, the surface that represents approximately the overall shape of a molecule as determined by the furthest extent of its electron cloud. On the other hand, if we make a plot where the value of electron density is relatively high, the resulting surface is one that approximately represents the region of covalent bonding in a molecule. Surfaces of low and high electron density are shown in this box for dimethyl ether. Similar models are shown for methane and ethane in Figs. 1.17 and 1.19.

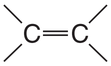


Dimethyl ether

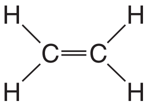
1.13 *The Structure of Ethene (Ethylene): sp² Hybridization*

The carbon atoms of many of the molecules that we have considered so far have used their four valence electrons to form four single covalent (sigma) bonds to four other atoms. We find, however, that many important organic compounds exist in which carbon atoms share more than two electrons with another atom. In molecules of these compounds some bonds that are formed are multiple covalent bonds. When two carbon atoms share two pairs of electrons, for example, the result is a carbon–carbon double bond:

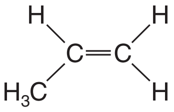
or



Hydrocarbons whose molecules contain a carbon–carbon double bond are called **alkenes**. Ethene (C_2H_4) and propene (C_3H_6) are both alkenes. (Ethene is also called ethylene, and propene is sometimes called propylene.)



Ethene



Propene

In ethene the only carbon–carbon bond is a double bond. Propene has one carbon–carbon single bond and one carbon–carbon double bond.

The spatial arrangement of the atoms of alkenes is different from that of alkanes. The six atoms of ethene are coplanar, and the arrangement of atoms around each carbon atom is triangular (Fig. 1.20).

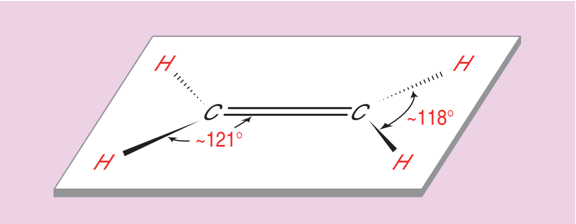


Figure 1.20 The structure and bond angles of ethene. The plane of the atoms is perpendicular to the paper. The dashed wedge bonds project behind the plane of the paper, and the solid wedge bonds project in front of the paper.

- Carbon–carbon double bonds are comprised of sp^2-hybridized carbon atoms.

The mathematical mixing of orbitals that furnish the **sp^2 orbitals** for our model can be visualized in the way shown in Fig. 1.21. The 2*s* orbital is mathematically mixed (or hybridized) with two of the 2*p* orbitals. (The hybridization procedure applies only to the orbitals, not to the electrons.) One 2*p* orbital is left unhybridized. One electron is then placed in each of the sp^2 hybrid orbitals and one electron remains in the 2*p* orbital.

The three sp^2 orbitals that result from hybridization are directed toward the corners of a regular triangle (with angles of 120° between them). The carbon *p* orbital that is not

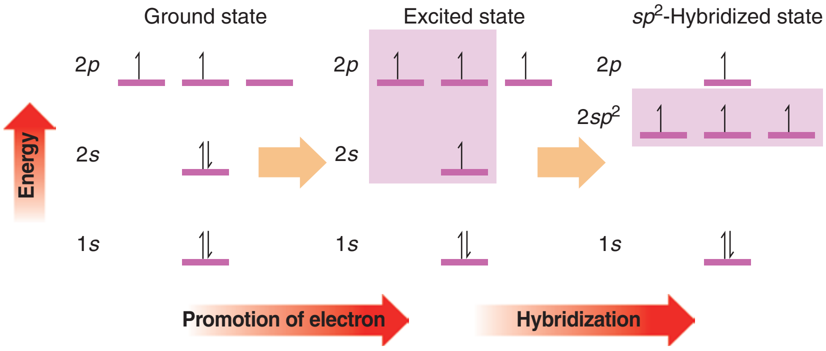


Figure 1.21 A process for deriving sp^2-hybridized carbon atoms.

hybridized is perpendicular to the plane of the triangle formed by the hybrid sp^2 orbitals (Fig. 1.22).

In our model for ethene (Fig. 1.23) we see the following:

- Two sp^2-hybridized carbon atoms form a sigma (σ) bond between them by overlap of one sp^2 orbital from each carbon. The remaining carbon sp^2 orbitals form σ bonds to four hydrogens through overlap with the hydrogen 1s orbitals. These five σ bonds account for 10 of the 12 valence electrons contributed by the two carbons and four hydrogens, and comprise the **σ-bond framework** of the molecule.

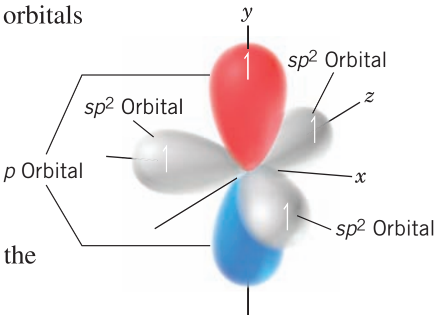


Figure 1.22 An sp^2-hybridized carbon atom.

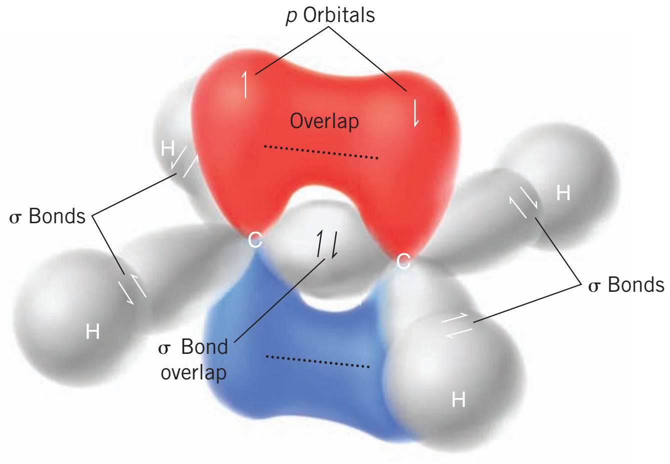
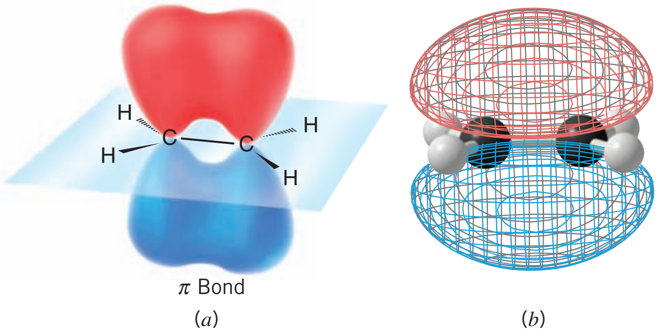


Figure 1.23 A model for the bonding molecular orbitals of ethene formed from two sp^2-hybridized carbon atoms and four hydrogen atoms.

- The remaining two bonding electrons are each located in an unhybridized p orbital of each carbon. Sideways overlap of these p orbitals and sharing of the two electrons between the carbons leads to a **pi (π) bond**. The overlap of these orbitals is shown schematically in Fig. 1.24.



(*a*) (*b*)

Figure 1.24 *(a)* A wedge–dashed wedge formula for the sigma bonds in ethene and a schematic depiction of the overlapping of adjacent p orbitals that form the π bond. *(b)* A calculated structure for ethene. The blue and red colors indicate opposite phase signs in each lobe of the π molecular orbital. A ball-and-stick model for the σ bonds in ethene can be seen through the mesh that indicates the π bond.

The bond angles that we would predict on the basis of sp^2-hybridized carbon atoms (120° all around) are quite close to the bond angles that are actually found (Fig. 1.20).

We can better visualize how these p orbitals interact with each other if we view a structure showing calculated molecular orbitals for ethene (Fig. 1.24). We see that the parallel p orbitals *overlap above and below the plane of the σ framework*.

Note the difference in shape of the bonding molecular orbital of a π bond as contrasted to that of a σ bond. A σ bond has cylindrical symmetry about a line connecting the two bonded nuclei. A π bond has a nodal plane passing through the two bonded nuclei and between the π molecular orbital lobes.

- When two p atomic orbitals combine to form a π bond, two molecular orbitals form: One is a bonding molecular orbital and the other is an antibonding molecular orbital.

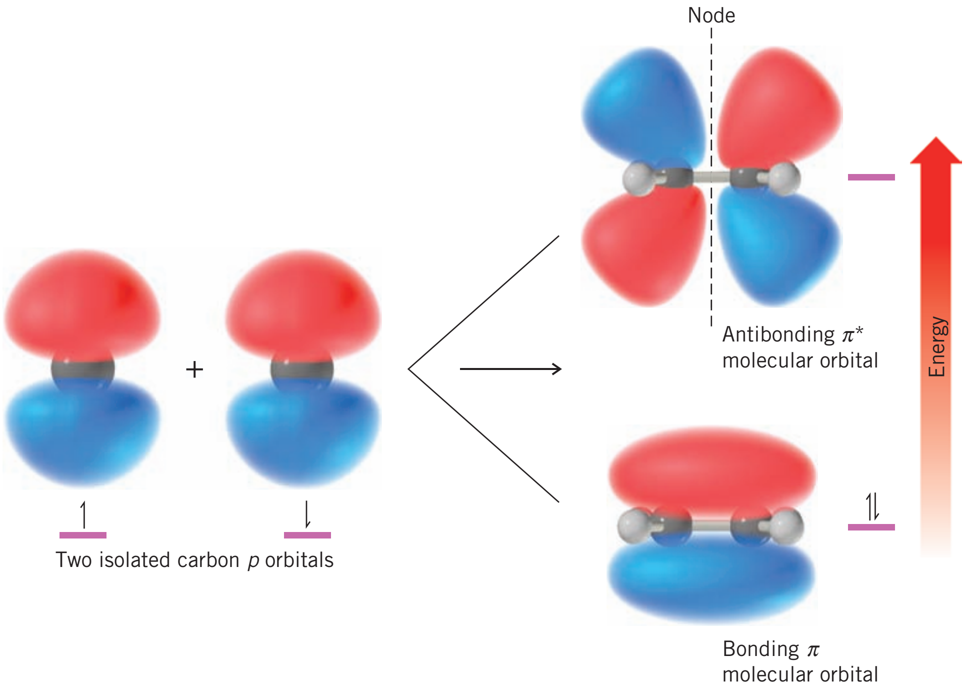
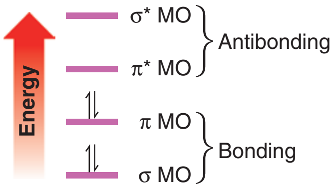


Figure 1.25 How two isolated carbon *p* orbitals combine to form two π (pi) molecular orbitals. The bonding MO is of lower energy. The higher energy antibonding MO contains an additional node. (Both orbitals have a node in the plane containing the C and H atoms.)

The bonding π molecular orbital results when *p*-orbital lobes of like signs overlap; the antibonding π molecular orbital results when opposite signs overlap (Fig. 1.25).

The bonding π orbital is the lower energy orbital and contains both π electrons (with opposite spins) in the ground state of the molecule. The region of greatest probability of finding the electrons in the bonding π orbital is a region generally situated above and below the plane of the σ-bond framework between the two carbon atoms. The antibonding π^* orbital is of higher energy, and it is not occupied by electrons when the molecule is in the ground state. It can become occupied, however, if the molecule absorbs light of the right frequency and an electron is promoted from the lower energy level to the higher one. The antibonding π^* orbital has a nodal plane between the two carbon atoms.

- To summarize, a carbon–carbon double bond consists of one σ bond and one π bond.



The relative energies of electrons involved in σ and π bonds.

The σ bond results from two sp^2 orbitals overlapping end to end and is symmetrical about an axis linking the two carbon atoms. The π bond results from a sideways overlap of two *p* orbitals; it has a nodal plane like a *p* orbital. In the ground state the electrons of the π bond are located between the two carbon atoms but generally above and below the plane of the σ-bond framework.

Electrons of the π bond have greater energy than electrons of the σ bond. The relative energies of the σ and π molecular orbitals (with the electrons in the ground state) are shown in the margin diagram. (The σ^* orbital is the antibonding sigma orbital.)

1.13A Restricted Rotation and the Double Bond

The σ–π model for the carbon–carbon double bond also accounts for an important property of the double bond:

- There is a large energy barrier to rotation associated with groups joined by a double bond.

Maximum overlap between the *p* orbitals of a π bond occurs when the axes of the *p* orbitals are exactly parallel. Rotating one carbon of the double bond 90° (Fig. 1.26) breaks the π bond, for then the axes of the *p* orbitals are perpendicular and there is no net overlap between

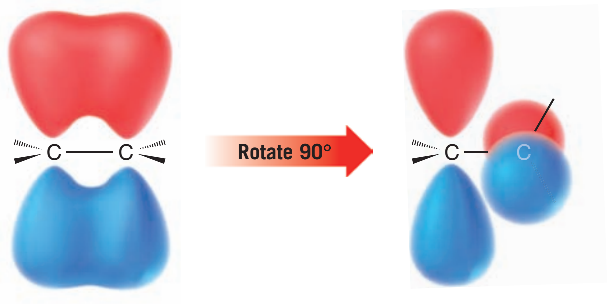
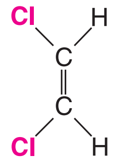


Figure 1.26 A stylized depiction of how rotation of a carbon atom of a double bond through an angle of 90° results in breaking of the π bond.

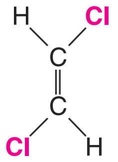
them. Estimates based on thermochemical calculations indicate that the strength of the π bond is 264 kJ mol^{-1}. This, then, is the barrier to rotation of the double bond. It is markedly higher than the rotational barrier of groups joined by carbon–carbon single bonds (13–26 kJ mol^{-1}). While groups joined by single bonds rotate relatively freely at room temperature, those joined by double bonds do not.

1.13B Cis–Trans Isomerism

Restricted rotation of groups joined by a double bond causes a new type of isomerism that we illustrate with the two dichloroethenes written as the following structures:



***cis*-1,2-Dichloroethene**



***trans*-1,2-Dichloroethene**

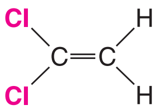
- These two compounds are isomers; they are different compounds that have the same molecular formula.

We can tell that they are different compounds by trying to place a model of one compound on a model of the other so that all parts coincide, that is, to try to **superpose** one on the other. We find that it cannot be done. Had one been **superposable** on the other, all parts of one model would correspond in three dimensions exactly with the other model. (*The notion of superposition is different from simply superimposing one thing on another*. The latter means only to lay one on the other without the necessary condition that all parts coincide.)

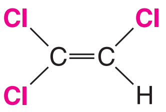
- We indicate that they are different isomers by attaching the prefix cis or trans to their names (*cis*, Latin: on this side; *trans*, Latin: across).

cis-1,2-Dichloroethene and *trans*-1,2-dichloroethene are not constitutional isomers because the connectivity of the atoms is the same in each. The two compounds ***differ only in the arrangement of their atoms in space***. Isomers of this kind are classified formally as **stereoisomers**, but often they are called simply cis–trans isomers. (We shall study stereoisomerism in detail in Chapters 4 and 5.)

The structural requirements for **cis–trans isomerism** will become clear if we consider a few additional examples. 1,1-Dichloroethene and 1,1,2-trichloroethene do not show this type of isomerism.

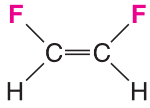


1,1-Dichloroethene
(no cis-trans isomerism)

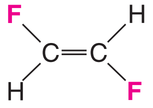


1,1,2-Trichloroethene
(no cis-trans isomerism)

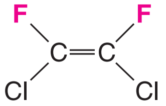
1,2-Difluoroethene and 1,2-dichloro-1,2-difluoroethene do exist as cis–trans isomers. Notice that we designate the isomer with two identical groups on the same side as being cis:



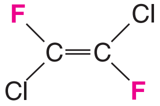
***cis*-1,2-Difluoroethene**



***trans*-1,2-Difluoroethene**



***cis*-1,2-Dichloro-1,2-difluoroethene**



***trans*-1,2-Dichloro-1,2-difluoroethene**

Clearly, then, ***cis–trans isomerism of this type is not possible if one carbon atom of the double bond bears two identical groups***.

Review Problem 1.13

Which of the following alkenes can exist as cis–trans isomers? Write their structures. Build handheld models to prove that one isomer is not superposable on the other.

(a) $CH_2{=}CHCH_2CH_3$

(b) $CH_3CH{=}CHCH_3$

(c) $CH_2{=}C(CH_3)_2$

(d) $CH_3CH_2CH{=}CHCl$

1.14 *The Structure of Ethyne (Acetylene): sp Hybridization*

Hydrocarbons in which two carbon atoms share three pairs of electrons between them, and are thus bonded by a triple bond, are called **alkynes**. The two simplest alkynes are ethyne and propyne.

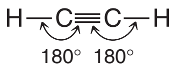
$H{-}C{\equiv}C{-}H$

Ethyne
(acetylene)
(C_2H_2)

$CH_3{-}C{\equiv}C{-}H$

Propyne
(C_3H_4)

Ethyne, a compound that is also called acetylene, consists of a linear arrangement of atoms. The $H{-}C{\equiv}C$ bond angles of ethyne molecules are 180°:



We can account for the structure of ethyne on the basis of orbital hybridization as we did for ethane and ethene. In our model for ethane (Section 1.12B) we saw that the carbon orbitals are sp^3 hybridized, and in our model for ethene (Section 1.13) we saw that they are sp^2 hybridized. In our model for ethyne we shall see that the carbon atoms are *sp hybridized*.

The mathematical process for obtaining the *sp* hybrid orbitals of ethyne can be visualized in the following way (Fig. 1.27).

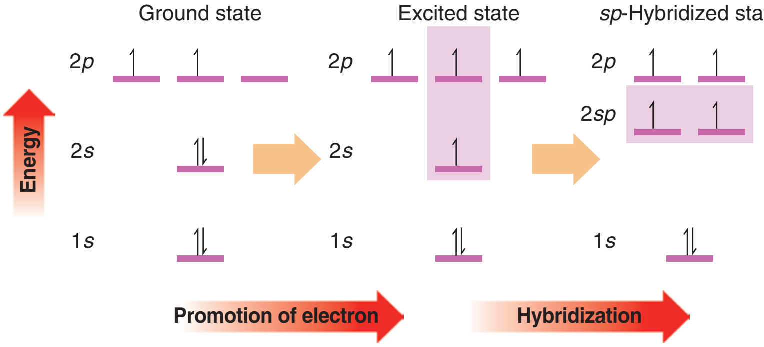


Figure 1.27 A process for deriving *sp*-hybridized carbon atoms.

- The 2*s* orbital and one 2*p* orbital of carbon are hybridized to form two ***sp* orbitals**.

The remaining two 2*p* orbitals are not hybridized.

Calculations show that the *sp* hybrid orbitals have their large positive lobes oriented at an angle of 180° with respect to each other. The two 2*p* orbitals that were not hybridized are each perpendicular to the axis that passes through the center of the two *sp* orbitals (Fig. 1.28). We place one electron in each orbital.

We envision the bonding molecular orbitals of ethyne being formed in the following way (Fig. 1.29).

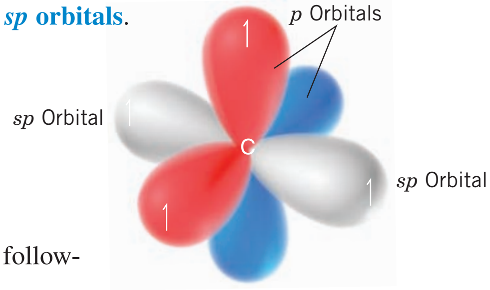


Figure 1.28 An *sp*-hybridized carbon atom.

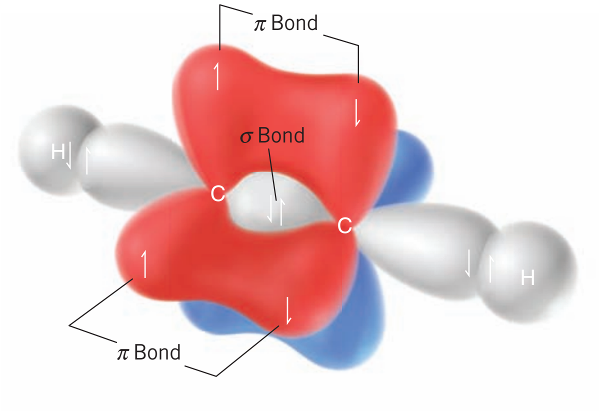


Figure 1.29 Formation of the bonding molecular orbitals of ethyne from two *sp*-hybridized carbon atoms and two hydrogen atoms. (Antibonding orbitals are formed as well, but these have been omitted for simplicity.)

- Two carbon atoms overlap *sp* orbitals to form a sigma bond between them (this is one bond of the triple bond). The remaining two *sp* orbitals at each carbon atom overlap with *s* orbitals from hydrogen atoms to produce two sigma C—H bonds.
- The two *p* orbitals on each carbon atom also overlap side to side to form two π bonds. These are the other two bonds of the triple bond.
- The carbon–carbon triple bond consists of two π bonds and one σ bond.

Structures for ethyne based on calculated molecular orbitals and electron density are shown in Fig. 1.30. Circular symmetry exists along the length of a triple bond (Fig. 1.30*b*). As a result, there is no restriction of rotation for groups joined by a triple bond (as compared with alkenes), and if rotation would occur, no new compound would form.

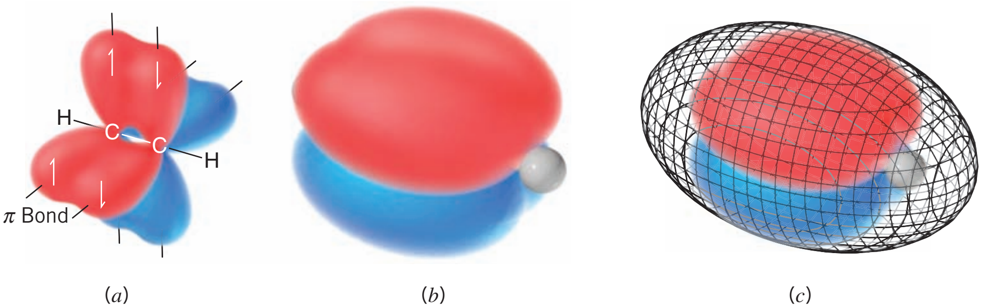


Figure 1.30 *(a)* The structure of ethyne (acetylene) showing the sigma-bond framework and a schematic depiction of the two pairs of *p* orbitals that overlap to form the two π bonds in ethyne. *(b)* A structure of ethyne showing calculated π molecular orbitals. Two pairs of π molecular orbital lobes are present, one pair for each π bond. The red and blue lobes in each π bond represent opposite phase signs. The hydrogen atoms of ethyne (white spheres) can be seen at each end of the structure (the carbon atoms are hidden by the molecular orbitals). *(c)* The mesh surface in this structure represents approximately the furthest extent of overall electron density in ethyne. Note that the overall electron density (but not the π-bonding electrons) extends over both hydrogen atoms.

1.14A Bond Lengths of Ethyne, Ethene, and Ethane

The carbon–carbon triple bond of ethyne is shorter than the carbon–carbon double bond of ethene, which in turn is shorter than the carbon–carbon single bond of ethane. The reason is that bond lengths are affected by the hybridization states of the carbon atoms involved.

- The greater the *s* orbital character in one or both atoms, the shorter is the bond. This is because *s* orbitals are spherical and have more electron density closer to the nucleus than do *p* orbitals.
- The greater the *p* orbital character in one or both atoms, the longer is the bond. This is because *p* orbitals are lobe-shaped with electron density extending away from the nucleus.

In terms of hybrid orbitals, an *sp* hybrid orbital has 50% *s* character and 50% *p* character. An sp^2 hybrid orbital has 33% *s* character and 67% *p* character. An sp^3 hybrid orbital has 25% *s* character and 75% *p* character. The overall trend, therefore, is as follows:

- Bonds involving *sp* hybrids are shorter than those involving sp^2 hybrids, which are shorter than those involving sp^3 hybrids. This trend holds true for both C—C and C—H bonds.

The bond lengths and bond angles of ethyne, ethene, and ethane are summarized in Fig. 1.31.

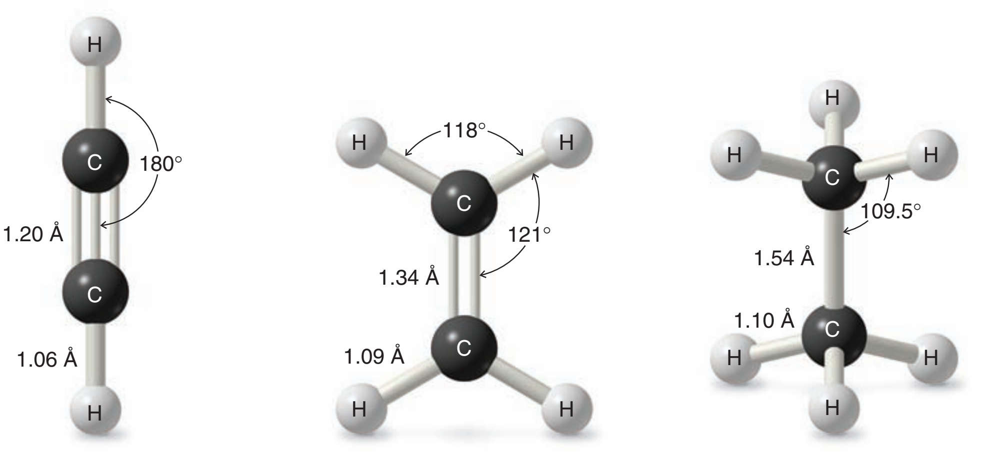
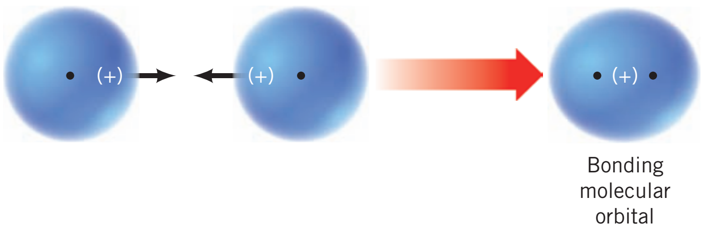


Figure 1.31 Bond angles and bond lengths of ethyne, ethene, and ethane.

1.15 *A Summary of Important Concepts That Come from Quantum Mechanics*

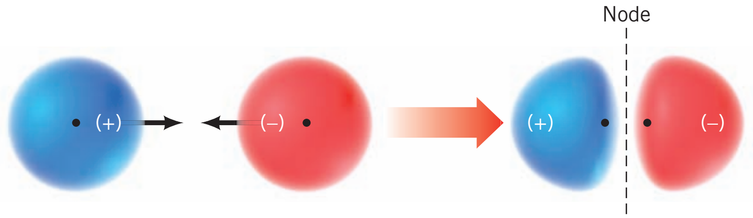
1. An **atomic orbital (AO)** corresponds to a region of space about the nucleus of a single atom where there is a high probability of finding an electron. Atomic orbitals called *s* orbitals are spherical; those called *p* orbitals are like two almost-tangent spheres. Orbitals can hold a maximum of two electrons when their spins are paired. Orbitals are described by the square of a wave function, ψ^2, and each orbital has a characteristic energy. The phase signs associated with an orbital may be + or −.
2. When atomic orbitals overlap, they combine to form **molecular orbitals (MOs)**. Molecular orbitals correspond to regions of space encompassing two (or more) nuclei where electrons are to be found. Like atomic orbitals, molecular orbitals can hold up to two electrons if their spins are paired.

3. When atomic orbitals with the same phase sign interact, they combine to form a **bonding molecular orbital**:



The electron probability density of a bonding molecular orbital is large in the region of space between the two nuclei where the negative electrons hold the positive nuclei together.

4. An **antibonding molecular orbital** forms when orbitals of opposite phase sign overlap:



An antibonding orbital has higher energy than a bonding orbital. The electron probability density of the region between the nuclei is small and it contains a **node**—a region where $\psi = 0$. Thus, having electrons in an antibonding orbital does not help hold the nuclei together. The internuclear repulsions tend to make them fly apart.

5. The **energy of electrons** in a bonding *molecular* orbital is less than the energy of the electrons in their separate *atomic* orbitals. The energy of electrons in an antibonding orbital is greater than that of electrons in their separate atomic orbitals.

6. The **number of molecular orbitals** always equals the number of atomic orbitals from which they are formed. Combining two atomic orbitals will always yield two molecular orbitals—one bonding and one antibonding.

7. **Hybrid atomic orbitals** are obtained by mixing (hybridizing) the wave functions for orbitals of different types (i.e., *s* and *p* orbitals) but from the same atom.

Helpful Hint

A summary of sp^3, sp^2, and sp hybrid orbital geometries.

8. Hybridizing three *p* orbitals with one *s* orbital yields four **sp^3 orbitals**. Atoms that are sp^3 hybridized direct the axes of their four sp^3 orbitals toward the corners of a tetrahedron. The carbon of methane is sp^3 hybridized and **tetrahedral**.

9. Hybridizing two *p* orbitals with one *s* orbital yields three **sp^2 orbitals**. Atoms that are sp^2 hybridized point the axes of their three sp^2 orbitals toward the corners of an equilateral triangle. The carbon atoms of ethene are sp^2 hybridized and **trigonal planar**.

10. Hybridizing one *p* orbital with one *s* orbital yields two ***sp* orbitals**. Atoms that are *sp* hybridized orient the axes of their two *sp* orbitals in opposite directions (at an angle of 180°). The carbon atoms of ethyne are *sp* hybridized and ethyne is a **linear** molecule.

11. A **sigma (σ) bond** (a type of single bond) is one in which the electron density has circular symmetry when viewed along the bond axis. In general, the skeletons of organic molecules are constructed of atoms linked by sigma bonds.

12. A **pi (π) bond**, part of double and triple carbon–carbon bonds, is one in which the electron densities of two adjacent parallel *p* orbitals overlap sideways to form a bonding pi molecular orbital.

1.16 *Molecular Geometry: The Valence Shell Electron Pair Repulsion Model*

We can predict the arrangement of atoms in molecules and ions on the basis of a relatively simple idea called the **valence shell electron pair repulsion (VSEPR) model**.

We apply the **VSEPR** model in the following way:

1. We consider molecules (or ions) in which the central atom is covalently bonded to two or more atoms or groups.
2. We consider all of the valence electron pairs of the central atom—both those that are shared in covalent bonds, called **bonding pairs**, and those that are unshared, called **nonbonding pairs** or **unshared pairs** or **lone pairs**.
3. Because electron pairs repel each other, the electron pairs of the valence shell tend to stay as far apart as possible. The repulsion between nonbonding pairs is generally greater than that between bonding pairs.
4. We arrive at the *geometry* of the molecule by considering all of the electron pairs, bonding and nonbonding, but we describe the *shape* of the molecule or ion by referring to the positions of the nuclei (or atoms) and not by the positions of the electron pairs.

Consider the following examples.

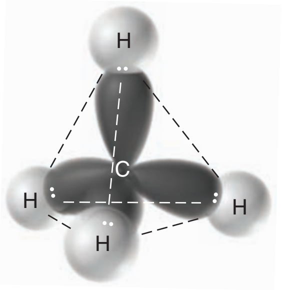


Figure 1.32 A tetrahedral shape for methane allows the maximum separation of the four bonding electron pairs.

1.16A Methane

The valence shell of methane contains four pairs of bonding electrons. Only a tetrahedral orientation will allow four pairs of electrons to have equal and maximum possible separation from each other (Fig. 1.32). Any other orientation, for example, a square planar arrangement, places some electron pairs closer together than others. Thus, methane has a tetrahedral shape.

The bond angles for any atom that has a regular tetrahedral structure are 109.5°. A representation of these angles in methane is shown in Fig. 1.33.

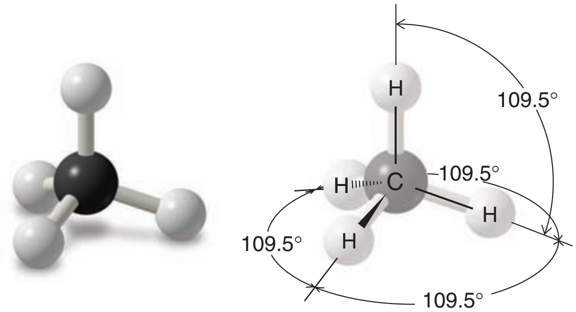


Figure 1.33 The bond angles of methane are 109.5°.

1.16B Ammonia

The shape of a molecule of ammonia (NH_3) is a **trigonal pyramid**. There are three bonding pairs of electrons and one nonbonding pair. The bond angles in a molecule of ammonia are 107°, a value very close to the tetrahedral angle (109.5°). We can write a general tetrahedral structure for the electron pairs of ammonia by placing the nonbonding pair at one corner (Fig. 1.34). A *tetrahedral arrangement* of the electron pairs explains the *trigonal pyramidal* arrangement of the four atoms. The bond angles are 107° (not 109.5°) because the nonbonding pair occupies more space than the bonding pairs.

Review Problem 1.14

What do the bond angles of ammonia suggest about the hybridization state of the nitrogen atom of ammonia?

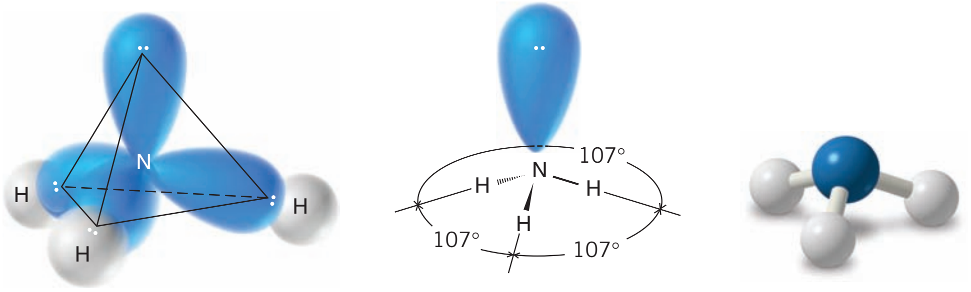


Figure 1.34 The tetrahedral arrangement of the electron pairs of an ammonia molecule that results when the nonbonding electron pair is considered to occupy one corner. This arrangement of electron pairs explains the trigonal pyramidal shape of the NH_3 molecule. Ball-and-stick models do not show unshared electrons.

1.16C Water

A molecule of water has an **angular** or **bent** shape. The H—O—H bond angle in a molecule of water is 104.5°, an angle that is also quite close to the 109.5° bond angles of methane.

We can write a general tetrahedral structure for the electron pairs of a molecule of water *if we place the two bonding pairs of electrons and the two nonbonding electron pairs at the corners of the tetrahedron*. Such a structure is shown in Fig. 1.35. A *tetrahedral arrangement* of the electron pairs accounts for the *angular arrangement* of the three atoms. The bond angle is less than 109.5° because the nonbonding pairs are effectively "larger" than the bonding pairs and, therefore, the structure is not perfectly tetrahedral.

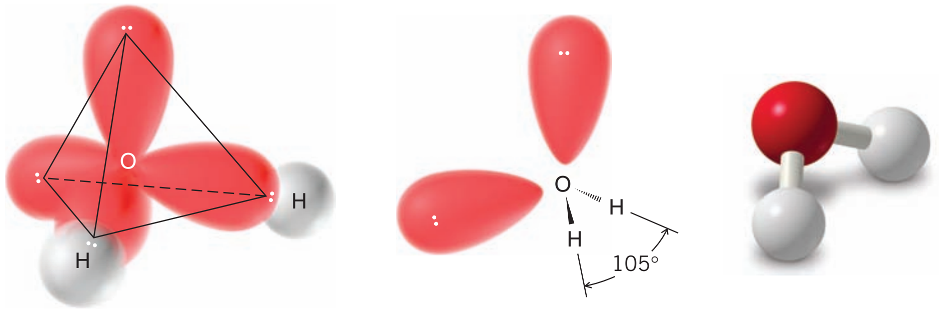


Figure 1.35 An approximately tetrahedral arrangement of the electron pairs of a molecule of water that results when the pairs of nonbonding electrons are considered to occupy corners. This arrangement accounts for the angular shape of the H_2O molecule.

Review Problem 1.15

What do the bond angles of water suggest about the hybridization state of the oxygen atom of water?

1.16D Boron Trifluoride

Boron, a group IIIA element, has only three valence electrons. In the compound boron trifluoride (BF_3) these three electrons are shared with three fluorine atoms. As a result, the boron atom in BF_3 has only six electrons (three bonding pairs) around it. Maximum separation of three bonding pairs occurs when they occupy the corners of an equilateral triangle. Consequently, in the boron trifluoride molecule the three fluorine atoms lie in a plane at the corners of an equilateral triangle (Fig. 1.36). Boron trifluoride is said to have a *trigonal planar structure*. The bond angles are 120°.

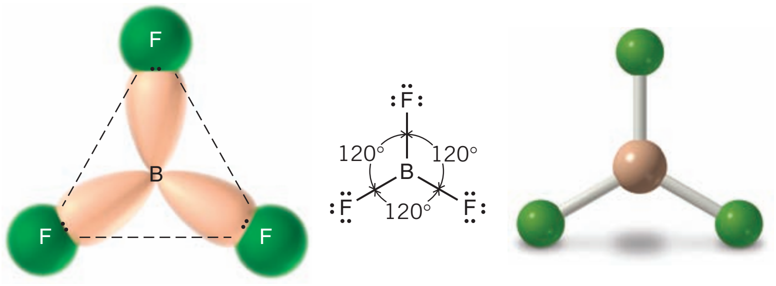


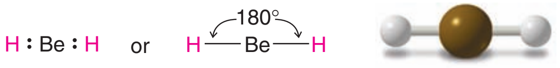
Figure 1.36 The triangular (trigonal planar) shape of boron trifluoride maximally separates the three bonding pairs.

Review Problem 1.16

What do the bond angles of boron trifluoride suggest about the hybridization state of the boron atom?

1.16E Beryllium Hydride

The central beryllium atom of BeH_2 has only two electron pairs around it; both electron pairs are bonding pairs. These two pairs are maximally separated when they are on opposite sides of the central atom, as shown in the following structures. This arrangement of the electron pairs accounts for the *linear geometry* of the BeH_2 molecule and its bond angle of 180°.



Linear geometry of BeH_2

Review Problem 1.17

What do the bond angles of beryllium hydride suggest about the hybridization state of the beryllium atom?

Review Problem 1.18

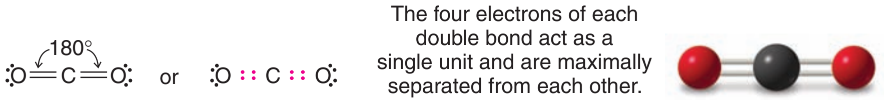
Use VSEPR theory to predict the geometry of each of the following molecules and ions:

(a) BH_4^-
(b) BeF_2
(c) NH_4^+
(d) H_2S
(e) BH_3
(f) CF_4
(g) SiF_4
(h) $:CCl_3^-$

1.16F Carbon Dioxide

The VSEPR method can also be used to predict the shapes of molecules containing multiple bonds if we assume that *all of the electrons of a multiple bond act as though they were a single unit* and, therefore, are located in the region of space between the two atoms joined by a multiple bond.

This principle can be illustrated with the structure of a molecule of carbon dioxide (CO_2). The central carbon atom of carbon dioxide is bonded to each oxygen atom by a double bond. Carbon dioxide is known to have a linear shape; the bond angle is 180°.



The four electrons of each double bond act as a single unit and are maximally separated from each other.

Such a structure is consistent with a maximum separation of the two groups of four bonding electrons. (The nonbonding pairs associated with the oxygen atoms have no effect on the shape.)

Review Problem 1.19

Predict the bond angles of

(a) $F_2C{=}CF_2$
(b) $CH_3C{\equiv}CCH_3$
(c) $HC{\equiv}N$

The shapes of several simple molecules and ions as predicted by VSEPR theory are shown in Table 1.4. In this table we have also included the hybridization state of the central atom.

TABLE 1.4 Shapes of Molecules and Ions from VSEPR Theory

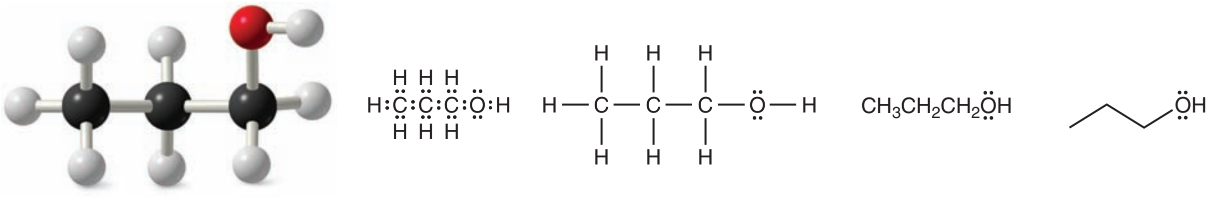
Number of Electron Pairs at Central Atom			Hybridization State of Central Atom	Shape of Molecule or Ion[a]	Examples
Bonding	**Nonbonding**	**Total**			
2	0	2	*sp*	Linear	BeH_2
3	0	3	sp^2	Trigonal planar	BF_3, CH_3^+
4	0	4	sp^3	Tetrahedral	CH_4, NH_4^+
3	1	4	$\sim sp^3$	Trigonal pyramidal	NH_3, CH_3^-
2	2	4	$\sim sp^3$	Angular	H_2O

[a] Referring to positions of atoms and excluding nonbonding pairs.

1.17 How to *Interpret and Write Structural Formulas*

Organic chemists use a variety of ways to write structural formulas. The most common types of representations are shown in Fig. 1.37 using propyl alcohol as an example. The **dot structure** shows all of the valence electrons, but writing it is tedious and time-consuming. The other representations are more convenient and are, therefore, more often used.

Sometimes we even omit unshared pairs when we write formulas. However, when we write chemical reactions, we see that it is necessary to include the unshared electron pairs when they participate in the reaction. It is a good idea, therefore, to get into the habit of writing the unshared (nonbonding) electron pairs in the structures you draw.

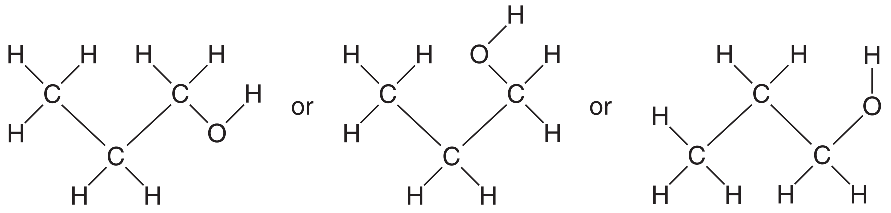


Ball-and-stick model (*a*)
Dot formula (*b*)
Dash formula (*c*)
Condensed formula (*d*)
Bond-line formula (*e*)

Figure 1.37 Structural formulas for propyl alcohol.

1.17A Dash Structural Formulas

If we look at the model for propyl alcohol given in Fig. 1.37*a* and compare it with the dot, dash, and condensed formulas in Figs. 1.37*b–d* we find that the chain of atoms is straight in those formulas. In the model, which corresponds more accurately to the actual shape of the molecule, the chain of atoms is not at all straight. Also of importance is this: ***Atoms joined by single bonds can rotate relatively freely with respect to one another***. (We discussed this point briefly in Section 1.12B.) This relatively free rotation means that the chain of atoms in propyl alcohol can assume a variety of arrangements like these:



Equivalent dash formulas for propyl alcohol

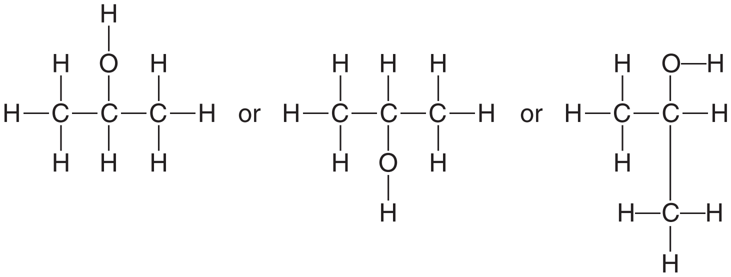
It also means that all of the structural formulas above are *equivalent* and all represent propyl alcohol. **Dash structural formulas** such as these indicate the way in which the atoms are attached to each other and *are not* representations of the actual shapes of the molecule.

Helpful Hint

It is important that you be able to recognize when a set of structural formulas has the same connectivity versus when they are constitutional isomers.

(Propyl alcohol does not have 90° bond angles. It has tetrahedral bond angles.) Dash structural formulas show what is called the **connectivity** of the atoms. *Constitutional isomers (Section 1.3A) have different connectivities and, therefore, must have different structural formulas.*

Consider the compound called isopropyl alcohol, whose formula we might write in a variety of ways:



Equivalent dash formulas for isopropyl alcohol

Isopropyl alcohol is a constitutional isomer (Section 1.3A) of propyl alcohol because its atoms are connected in a different order and both compounds have the same molecular formula, C_3H_8O. In isopropyl alcohol the OH group is attached to the central carbon; in propyl alcohol it is attached to an end carbon.

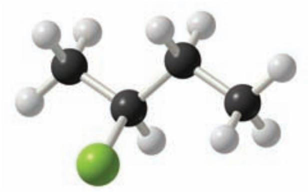
- In problems you will often be asked to write structural formulas for all the isomers that have a given molecular formula. Do not make the error of writing several equivalent formulas, like those that we have just shown, mistaking them for different constitutional isomers.

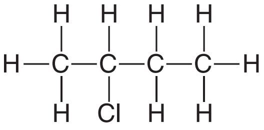
Review Problem 1.20

There are actually three constitutional isomers with the molecular formula C_3H_8O. We have seen two of them in propyl alcohol and isopropyl alcohol. Write a dash formula for the third isomer.

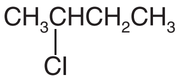
1.17B Condensed Structural Formulas

Condensed structural formulas are somewhat faster to write than dash formulas and, when we become familiar with them, they will impart all the information that is contained in the dash structure. In condensed formulas all of the hydrogen atoms that are attached to a particular carbon are usually written immediately after the carbon. In fully condensed formulas, all of the atoms that are attached to the carbon are usually written immediately after that carbon, listing hydrogens first. For example,





Dash formula

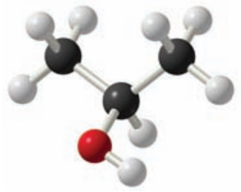


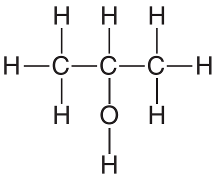
$$CH_3\underset{\displaystyle Cl}{\underset{|}{C}}HCH_2CH_3$$

or $CH_3CHClCH_2CH_3$

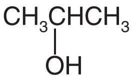
Condensed formulas

The condensed formula for isopropyl alcohol can be written in four different ways:





Dash formula



$$CH_3\underset{\displaystyle OH}{\underset{|}{C}}HCH_3$$

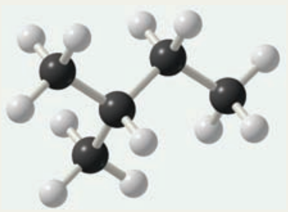
or $CH_3CH(OH)CH_3$

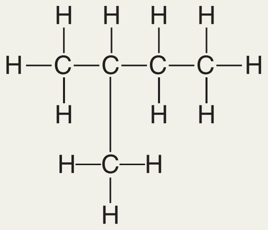
$CH_3CHOHCH_3$ or $(CH_3)_2CHOH$

Condensed formulas

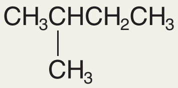
Solved Problem 1.8

Write a condensed structural formula for the compound that follows:



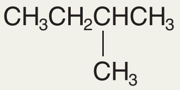


ANSWER



$$\underset{\displaystyle CH_3}{CH_3\overset{}{C}HCH_2CH_3}$$ or $CH_3CH(CH_3)CH_2CH_3$ or $(CH_3)_2CHCH_2CH_3$

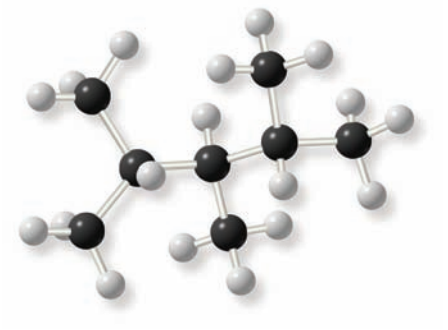
or $CH_3CH_2CH(CH_3)_2$ or



$$CH_3CH_2\underset{\displaystyle CH_3}{CH}CH_3$$

Review Problem 1.21

Write a condensed structural formula for the following compound.



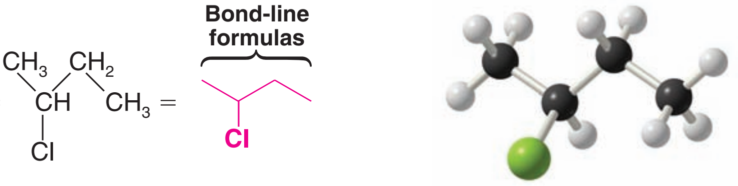
1.17C Bond-Line Formulas

The most common type of structural formula used by organic chemists, and the fastest to draw, is the **bond-line formula**. The formula in Fig. 1.37*e* is a bond-line formula for propyl alcohol. The sooner you master the use of bond-line formulas, the more quickly you will be able to draw molecules when you take notes and work problems. And, lacking all of the symbols that are explicitly shown in dash and condensed structural formulas, bond-line formulas allow you to more quickly interpret molecular connectivity and compare one molecular formula with another.

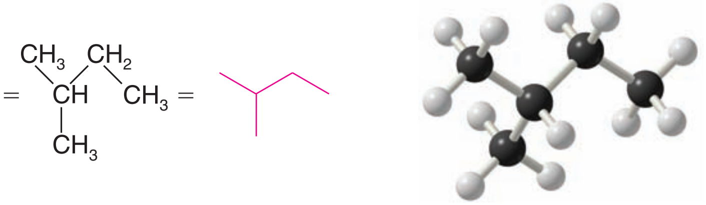
The efficiencies of drawing bond-line formulas come from the fact that no Cs are written for carbon atoms, and generally no Hs are shown for hydrogen atoms, unless they are needed to give a three-dimensional perspective to the molecule (and in that case we use solid or dashed wedges for bonds to the out-of-plane atoms, as described in the following section). Instead, in bond-line formulas ordinary lines represent bonds, and carbon atoms are inferred at each bend in the line and at the ends of lines.

The number of hydrogen atoms bonded to each carbon is also inferred, by assuming that as many hydrogen atoms are present as needed to fill the valence shell of each carbon, unless a charge is indicated. When an atom other than carbon is present, the symbol for that element is written in the formula at the appropriate location, i.e., in place of a bend or at the terminus of the line leading to the atom. Hydrogen atoms bonded to atoms other than carbon (e.g., oxygen or nitrogen) are written explicitly. And, as mentioned above, hydrogen atoms are shown where needed to help specify three dimensions using solid or dashed wedges. Consider the following examples of molecules depicted by bond-line formulas.

$CH_3CHClCH_2CH_3$ =



$CH_3CH(CH_3)CH_2CH_3$ =

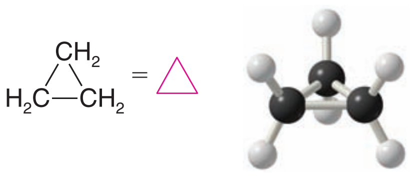


$(CH_3)_2NCH_2CH_3$ =

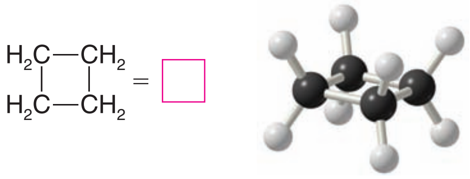
Helpful Hint

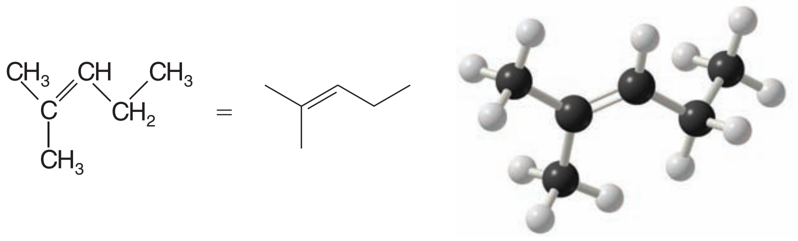
As you become more familiar with organic molecules, you will find bond-line formulas to be very useful tools for representing structures.

Bond-line formulas are easy to draw for molecules with multiple bonds and for cyclic molecules, as well. The following are some examples.

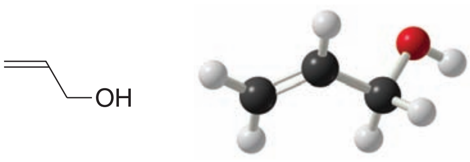


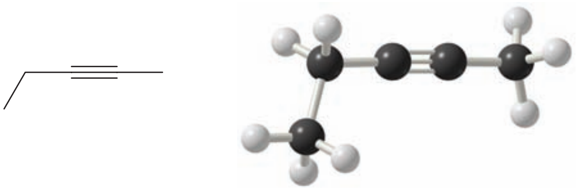
and





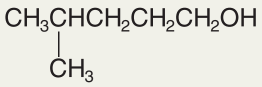
$CH_2{=}CHCH_2OH$ =





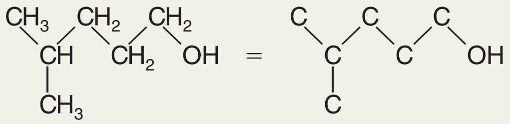
Solved Problem 1.9

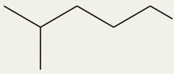
Write the bond-line formula for



$$\underset{\displaystyle CH_3}{CH_3\underset{|}{C}HCH_2CH_2CH_2OH}$$

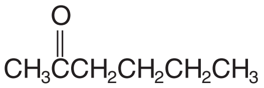
STRATEGY AND ANSWER First, for the sake of practice, we outline the carbon skeleton, including the OH group, as follows:



Then we write the bond-line formula as  OH. As you gain experience you will likely skip the intermediate steps shown above and proceed directly to writing bond-line formulas.

Review Problem 1.22

Write each of the following condensed structural formulas as a bond-line formula:

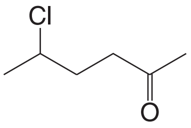
(a) $(CH_3)_2CHCH_2CH_3$
(b) $(CH_3)_2CHCH_2CH_2OH$
(c) $(CH_3)_2C{=}CHCH_2CH_3$
(d) $CH_3CH_2CH_2CH_2CH_3$
(e) $CH_3CH_2CH(OH)CH_2CH_3$
(f) $CH_2{=}C(CH_2CH_3)_2$
(g)  $CH_3\overset{\displaystyle O}{\overset{\|}{C}}CH_2CH_2CH_2CH_3$
(h) $CH_3CHClCH_2CH(CH_3)_2$

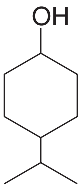
Review Problem 1.23

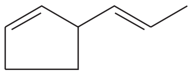
Which molecules in Review Problem 1.22 form sets of constitutional isomers?

Review Problem 1.24

Write a dash formula for each of the following bond-line formulas:

(a) 

(b) 

(c) 

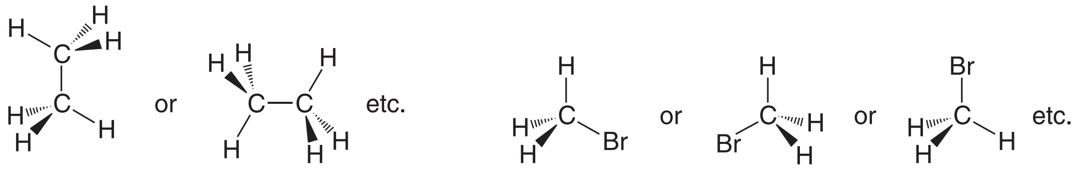
1.17D Three-Dimensional Formulas

None of the formulas that we have described so far conveys any information about how the atoms of a molecule are arranged in space. There are several types of representations that do this. The types of formulas that we shall use most of the time are shown in Fig. 1.38. In these representations, bonds that project upward out of the plane of the paper are indicated by a solid wedge (◀), those that lie behind the plane are indicated with a dashed wedge (⋯), and those bonds that lie in the plane of the page are indicated by a line (—). For tetrahedral atoms, notice that we draw the two bonds that are in the plane of the page with an angle of approximately 109° between them and that proper three-dimensional perspective then requires the wedge and dashed-wedge bonds to be drawn near each other on the page (i.e., the atom in front nearly eclipses the atom behind). We can draw trigonal planar atoms either with all bonds in the plane of the page separated by approximately 120° or with one of the three bonds in the plane of the page, one behind, and one in front (as in Fig. 1.20). Atoms with linear bonding geometry are best drawn with all bonds to those atoms in the plane of the page. Lastly, when drawing three-dimensional formulas is it generally best to draw as many carbon atoms as possible in the plane of the paper, allowing substituent groups or hydrogen atoms to be primarily those for which wedge or dashed-wedge bonds

Helpful Hint

Wedge and dashed-wedge formulas are a tool for unambiguously showing three dimensions.

are used. Note that when using bond-line formulas we continue to omit hydrogen atoms unless they are relevant to clarifying the three-dimensional perspective of some other group. Figure 1.38 shows some examples of three-dimensional formulas.



Ethane **Bromomethane**

Examples of bond-line formulas that include three-dimensional representations

An example involving trigonal planar geometry

An example involving linear geometry

Figure 1.38 Some examples of three-dimensional formulas.

Review Problem 1.25

Write three-dimensional (wedge–dashed wedge–line) representations for each of the following:

(a) CH_3Cl **(b)** CH_2Cl_2 **(c)** CH_2BrCl **(d)** CH_3CH_2Cl

---

1.18 *Applications of Basic Principles*

Throughout the early chapters of this book we review certain basic principles that underlie and explain much of the chemistry we shall be studying. Consider the following principles and how they apply in this chapter.

Opposite Charges Attract We see this principle operating in our explanations for covalent and ionic bonds (Sections 1.11 and 1.4A). It is the attraction of the *positively* charged nuclei for the *negatively* charged electrons that underlies our explanation for the covalent bond. It is the attraction of the oppositely charged ions in crystals that explains the ionic bond.

Like Charges Repel It is the repulsion of the electrons in covalent bonds of the valence shell of a molecule that is central to the valence shell electron pair repulsion model for explaining molecular geometry. And, although it is not so obvious, this same factor underlies the explanations of molecular geometry that come from orbital hybridization because these repulsions are taken into account in calculating the orientations of the hybrid orbitals.

Nature Tends toward States of Lower Potential Energy This principle explains so much of the world around us. It explains why water flows downhill: The potential energy of the water at the bottom of the hill is lower than that at the top. (We say that water is in a more stable state at the bottom.) This principle underlies the aufbau principle (Section 1.10A): In its lowest energy state, the electrons of an atom occupy the lowest energy orbitals available [but Hund's rule still applies, as well as the Pauli exclusion principle (Section 1.10A), allowing only two electrons per orbital]. Similarly in molecular orbital theory (Section 1.11), electrons fill lower energy bonding molecular orbitals first because this gives the molecule lower potential energy (or greater stability). Energy has to be provided to move an electron to a higher orbital and provide an excited (less stable) state (Review Problem 1.12).

Orbital Overlap Stabilizes Molecules This principle is part of our explanation for covalent bonds. When orbitals of the same phase from different nuclei overlap, the electrons in these orbitals can be shared by both nuclei, resulting in stabilization. The result is a covalent bond.

In This Chapter

In Chapter 1 you have studied concepts and skills that are absolutely essential to your success in organic chemistry. You should now be able to use the periodic table to determine the number of valence electrons an atom has in its neutral state or as an ion. You should be able to use the periodic table to compare the relative electronegativity of one element with another, and determine the formal charge of an atom or ion. Electronegativity and formal charge are key concepts in organic chemistry.

You should be able to draw chemical formulas that show all of the valence electrons in a molecule (Lewis structures), using lines for bonds and dots to show unshared electrons. You should be proficient in representing structures as dash structural formulas, condensed structural formulas, and bond-line structural formulas. In particular, the more quickly you become skilled at using and interpreting bond-line formulas, the faster you will be able to process structural information in organic chemistry. You have also learned about resonance structures, the use of which will help us in understanding a variety of concepts in later chapters.

Lastly, you have learned to predict the three-dimensional structure of molecules using the valence shell electron pair repulsion (VSEPR) model and molecular orbital (MO) theory. An ability to predict three-dimensional structure is critical to understanding the properties and reactivity of molecules.

We encourage you to do all of the problems that your instructor has assigned. We also recommend that you use the summary and review tools in each chapter, such as the concept map that follows. Concept maps can help you see the flow of concepts in a chapter and also help remind you of key points. In fact, we encourage you to build your own concept maps for review when the opportunity arises.

Work especially hard to solidify your knowledge from this and other early chapters in the book. These chapters have everything to do with helping you learn basic tools you need for success throughout organic chemistry.

Key Terms and Concepts

The key terms and concepts that are highlighted in **bold, blue text** within the chapter are defined in the glossary (at the back of the book) and have hyperlinked definitions in the accompanying ***WileyPLUS*** course (www.wileyplus.com).

Problems

Note to Instructors: Many of the homework problems are available for assignment via *WileyPLUS*, an online teaching and learning solution.

ELECTRON CONFIGURATION

1.26 Which of the following ions possess the electron configuration of a noble gas?

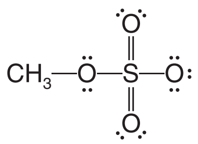
(a) Na^+	**(c)** F^+	**(e)** Ca^{2+}	**(g)** O^{2-}
(b) Cl^-	**(d)** H^-	**(f)** S^{2-}	**(h)** Br^+

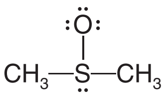
LEWIS STRUCTURES

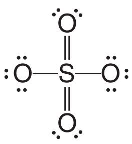
1.27 Write a Lewis structure for each of the following:

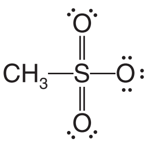
(a) $SOCl_2$	**(b)** $POCl_3$	**(c)** PCl_5	**(d)** $HONO_2$ (HNO_3)

1.28 Give the formal charge (if one exists) on each atom of the following:

(a)

$CH_3{-}\ddot{O}{-}\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{S}}{-}\ddot{O}:$

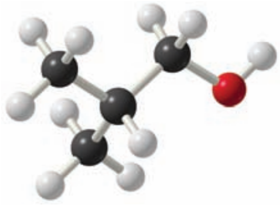
(b)

$CH_3{-}\overset{\overset{:\ddot{O}:}{|}}{\ddot{S}}{-}CH_3$

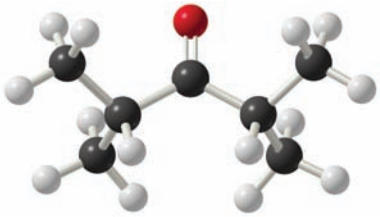
(c)

$:\ddot{O}{-}\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{S}}{-}\ddot{O}:$

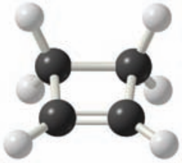
(d)

$CH_3{-}\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{S}}{-}\ddot{O}:$

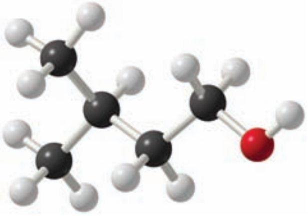
STRUCTURAL FORMULAS AND ISOMERISM

1.29 Write a condensed structural formula for each compound given here.


(a)

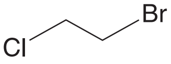
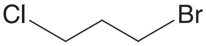

(b)

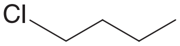

(c)

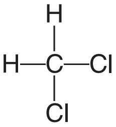
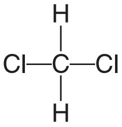

(d)

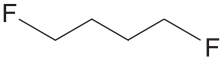
1.30 What is the molecular formula for each of the compounds given in Exercise 1.29?

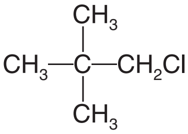
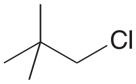
1.31 Consider each pair of structural formulas that follow and state whether the two formulas represent the same compound, whether they represent different compounds that are constitutional isomers of each other, or whether they represent different compounds that are not isomeric.

(a)  and 

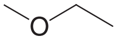
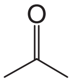
(b)  and $ClCH_2CH(CH_3)_2$

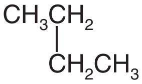
(c)  $H-\overset{H}{\underset{Cl}{C}}-Cl$ and  $Cl-\overset{H}{\underset{H}{C}}-Cl$

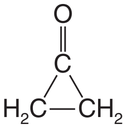
(d)  and $FCH_2CH_2CH_2CH_2F$

(e)  $CH_3-\overset{CH_3}{\underset{CH_3}{C}}-CH_2Cl$ and 

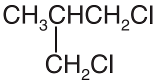
(f) $CH_2{=}CHCH_2CH_3$ and 

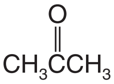
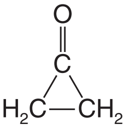
(g)  and 

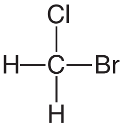
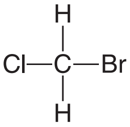
(h)  $CH_3\underset{CH_2CH_3}{CH_2}$ and 

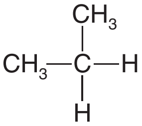
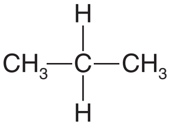
(i) $CH_3OCH_2CH_3$ and 

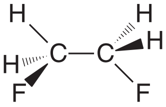
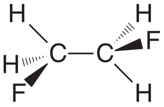
(j) $CH_2ClCHClCH_3$ and $CH_3CHClCH_2Cl$

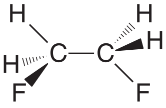
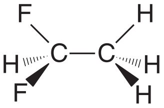
(k) $CH_3CH_2CHClCH_2Cl$ and  $CH_3\underset{CH_2Cl}{CH}CH_2Cl$

(l)  $CH_3\overset{O}{\overset{\|}{C}}CH_3$ and 

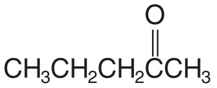
(m)  $H-\overset{Cl}{\underset{H}{C}}-Br$ and  $Cl-\overset{H}{\underset{H}{C}}-Br$

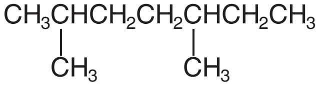
(n)  $CH_3-\overset{CH_3}{\underset{H}{C}}-H$ and  $CH_3-\overset{H}{\underset{H}{C}}-CH_3$

(o)  and 

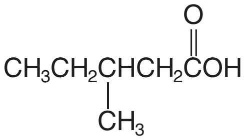
(p)  and 

1.32 Rewrite each of the following using bond-line formulas:

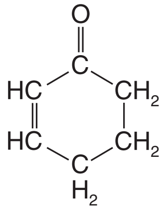
(a) 
$CH_3CH_2CH_2\overset{O}{\overset{\|}{C}}CH_3$

(b) 
$CH_3\underset{CH_3}{\underset{|}{C}H}CH_2CH_2\underset{CH_3}{\underset{|}{C}H}CH_2CH_3$

(c) $(CH_3)_3CCH_2CH_2CH_2OH$

(d) 
$CH_3CH_2\underset{CH_3}{\underset{|}{C}H}CH_2\overset{O}{\overset{\|}{C}}OH$

(e) $CH_2{=}CHCH_2CH_2CH{=}CHCH_3$

(f) 

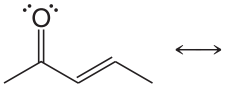
1.33 Write structural formulas of your choice for all of the constitutional isomers with the molecular formula C_4H_8.

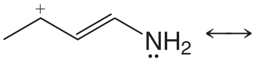
1.34 Write structural formulas for at least three constitutional isomers with the molecular formula CH_3NO_2. (In answering this question you should assign a formal charge to any atom that bears one.)

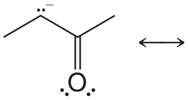
RESONANCE STRUCTURES

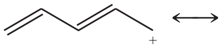
1.35 For the following write all possible resonance structures. Be sure to include formal charges where appropriate.

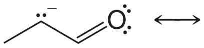
(a) 

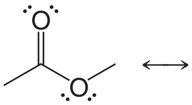
(b) 

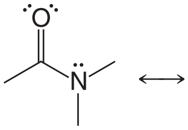
(c) 

(d) 

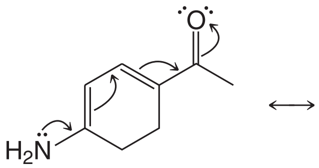
(e) 

(f) 

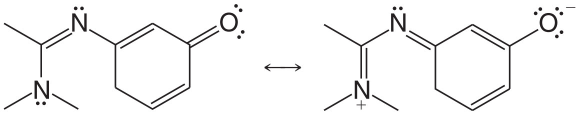
(g) 

(h) 

1.36 Write the resonance structure that would result from moving the electrons in the way indicated by the curved arrows.



1.37 Show the curved arrows that would convert **A** into **B**.



A **B**

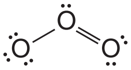
1.38 **(a)** Cyanic acid (H—O—C≡N) and isocyanic acid (H—N=C=O) differ in the positions of their electrons but their structures do not represent resonance structures. Explain. **(b)** Loss of a proton from cyanic acid yields the same anion as that obtained by loss of a proton from isocyanic acid. Explain.

1.39 Consider a chemical species (either a molecule or an ion) in which a carbon atom forms three single bonds to three hydrogen atoms and in which the carbon atom possesses no other valence electrons. **(a)** What formal charge would the carbon atom have? **(b)** What total charge would the species have? **(c)** What shape would you expect this species to have? **(d)** What would you expect the hybridization state of the carbon atom to be?

1.40 Consider a chemical species like the one in the previous problem in which a carbon atom forms three single bonds to three hydrogen atoms, but in which the carbon atom possesses an unshared electron pair. **(a)** What formal charge would the carbon atom have? **(b)** What total charge would the species have? **(c)** What shape would you expect this species to have? **(d)** What would you expect the hybridization state of the carbon atom to be?

1.41 Consider another chemical species like the ones in the previous problems in which a carbon atom forms three single bonds to three hydrogen atoms but in which the carbon atom possesses a single unpaired electron. **(a)** What formal charge would the carbon atom have? **(b)** What total charge would the species have? **(c)** Given that the shape of this species is trigonal planar, what would you expect the hybridization state of the carbon atom to be?

1.42 Ozone (O_3) is found in the upper atmosphere where it absorbs highly energetic ultraviolet (UV) radiation and thereby provides the surface of Earth with a protective screen (cf. Section 10.11E). One possible resonance structure for ozone is the following:



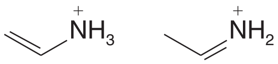
(a) Assign any necessary formal charges to the atoms in this structure. **(b)** Write another equivalent resonance structure for ozone. **(c)** What do these resonance structures predict about the relative lengths of the two oxygen–oxygen bonds of ozone? **(d)** In the structure above, and the one you have written, assume an angular shape for the ozone molecule. Is this shape consistent with VSEPR theory? Explain your answer.

1.43 Write resonance structures for the azide ion, N_3^-. Explain how these resonance structures account for the fact that both bonds of the azide ion have the same length.

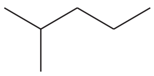
1.44 Write structural formulas of the type indicated: **(a)** bond-line formulas for seven constitutional isomers with the formula $C_4H_{10}O$; **(b)** condensed structural formulas for two constitutional isomers with the formula C_2H_7N; **(c)** condensed structural formulas for four constitutional isomers with the formula C_3H_9N; **(d)** bond-line formulas for three constitutional isomers with the formula C_5H_{12}.

1.45 What is the relationship between the members of the following pairs? That is, are they constitutional isomers, the same, or something else (specify)?

(a)

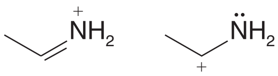


(b)

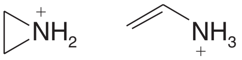


$CH_3CH_2CH_2CH(CH_3)_2$

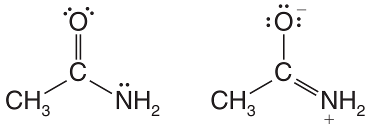
(c)



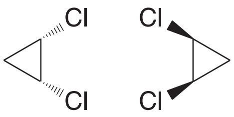
(d)



(e)



(f)



Challenge Problems

1.46 In Chapter 15 we shall learn how the nitronium ion, NO_2^+, forms when concentrated nitric and sulfuric acids are mixed. **(a)** Write a Lewis structure for the nitronium ion. **(b)** What geometry does VSEPR theory predict for the NO_2^+ ion? **(c)** Give a species that has the same number of electrons as NO_2^+.

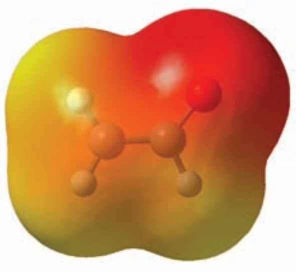
1.47 Given the following sets of atoms, write bond-line formulas for all of the possible constitutionally isomeric compounds or ions that could be made from them. Show all unshared electron pairs and all formal charges, if any.

Set	C atoms	H atoms	Other
A	3	6	2 Br atoms
B	3	9	1 N atom and 1 O atom (not on same C)
C	3	4	1 O atom
D	2	7	1 N atom and 1 proton
E	3	7	1 extra electron

1.48 Open computer molecular models for dimethyl ether, dimethylacetylene, and *cis*-1,2-dichloro-1,2-difluoroethene from the 3D Molecular Models section of the book's website. By interpreting the computer molecular model for each one, draw **(a)** a dash formula, **(b)** a bond-line formula, and **(c)** a three-dimensional dashed-wedge formula. Draw the models in whatever perspective is most convenient—generally the perspective in which the most atoms in the chain of a molecule can be in the plane of the paper.

1.49 Boron is a group IIIA element. Open the molecular model for boron trifluoride from the 3D Molecular Models section of the book's website. Near the boron atom, above and below the plane of the atoms in BF_3, are two relatively large lobes. Considering the position of boron in the periodic table and the three-dimensional and electronic structure of BF_3, what type of orbital does this lobe represent? Is it a hybridized orbital or not?

1.50 There are two contributing resonance structures for an anion called acetaldehyde enolate, whose condensed molecular formula is CH_2CHO^-. Draw the two resonance contributors and the resonance hybrid, then consider the map of electrostatic potential (MEP) shown below for this anion. Comment on whether the MEP is consistent or not with predominance of the resonance contributor you would have predicted to be represented most strongly in the hybrid.



Learning Group Problems

Consider the compound with the following condensed molecular formula:

$$CH_3CHOHCH{=}CH_2$$

Helpful Hint

Your instructor will tell you how to work these problems as a Learning Group.

1. Write a full dash structural formula for the compound.

2. Show all nonbonding electron pairs on your dash structural formula.

3. Indicate any formal charges that may be present in the molecule.

4. Label the hybridization state at every carbon atom and the oxygen.

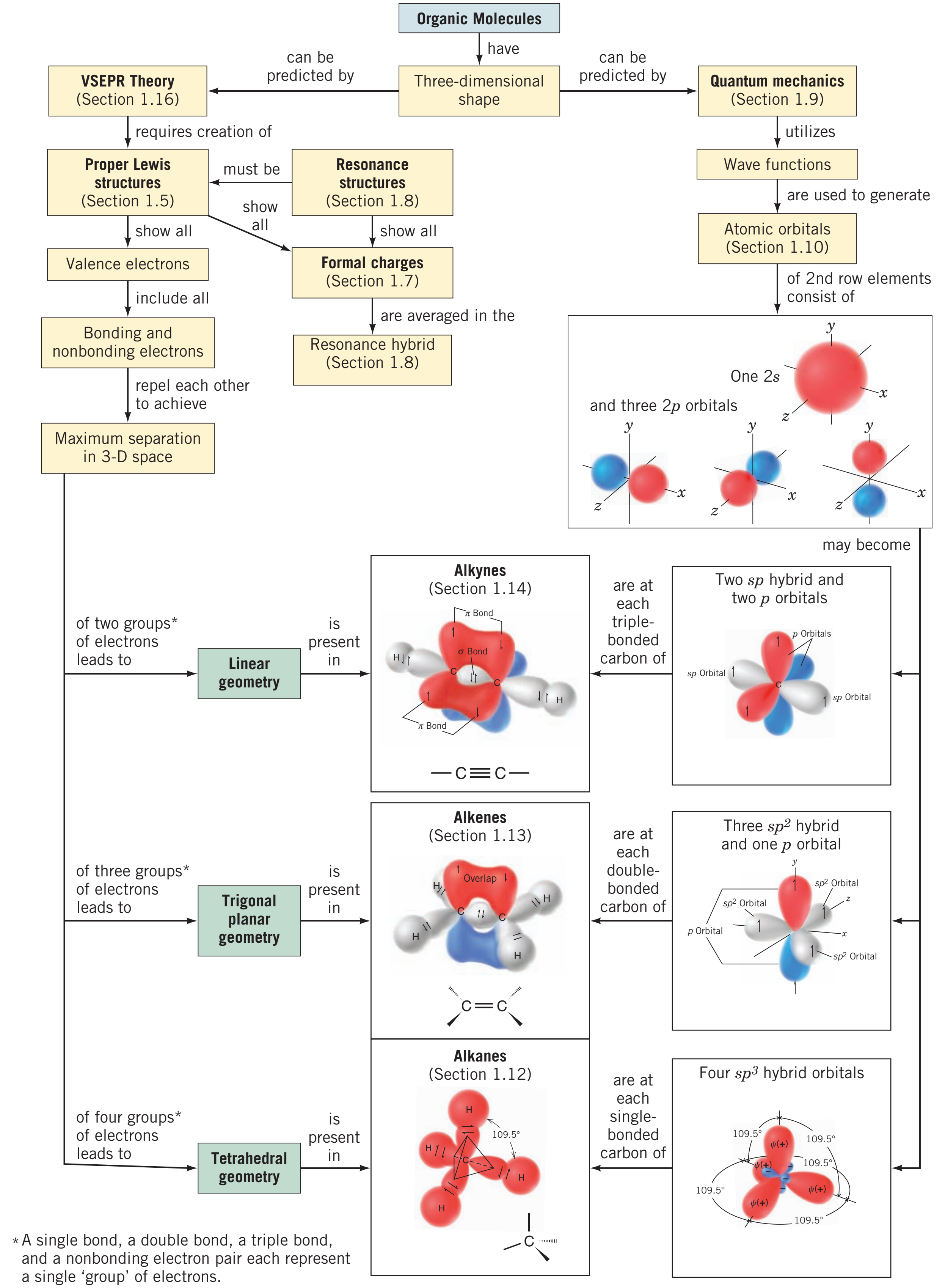
5. Draw a three-dimensional perspective representation for the compound showing approximate bond angles as clearly as possible. Use ordinary lines to indicate bonds in the plane of the paper, solid wedges for bonds in front of the paper, and dashed wedges for bonds behind the paper.

6. Label all the bond angles in your three-dimensional structure.

7. Draw a bond-line formula for the compound.

8. Devise two structures, each having two *sp*-hybridized carbons and the molecular formula C_4H_6O. Create one of these structures such that it is linear with respect to all carbon atoms. Repeat parts 1–7 above for both structures.

CONCEPT MAP

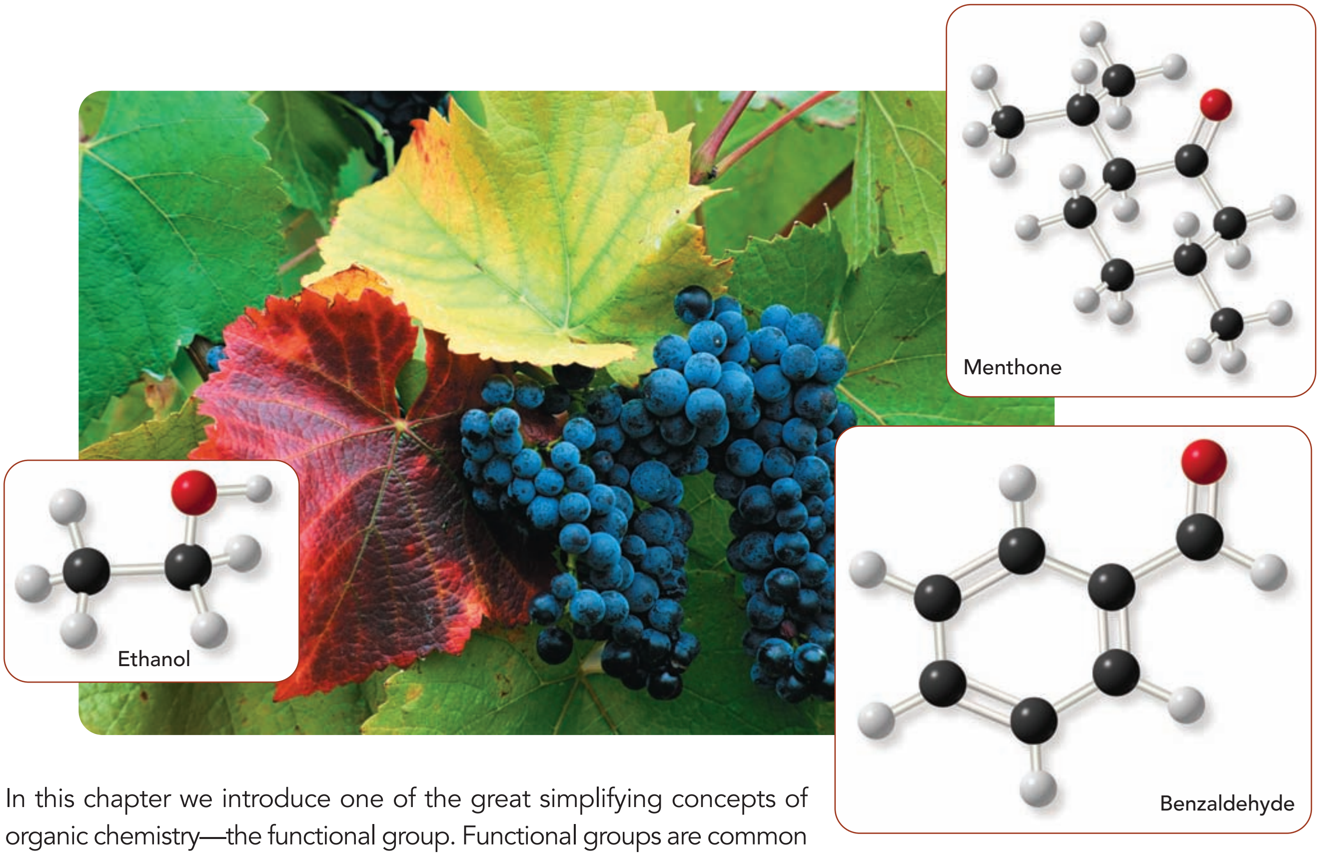


*A single bond, a double bond, a triple bond, and a nonbonding electron pair each represent a single 'group' of electrons.

2

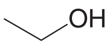
Families of Carbon Compounds

Functional Groups, Intermolecular Forces, and Infrared (IR) Spectroscopy

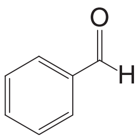


In this chapter we introduce one of the great simplifying concepts of organic chemistry—the functional group. Functional groups are common and specific arrangements of atoms that impart predictable reactivity and properties to a molecule. Even though there are millions of organic compounds, you may be relieved to know that we can readily understand much about whole families of compounds simply by learning about the properties of the common functional groups.

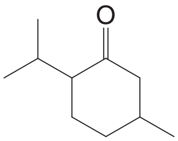
For example, all alcohols contain an —OH (hydroxyl) functional group attached to a saturated carbon bearing nothing else but carbon or hydrogen. Alcohols as simple as ethanol in alcoholic beverages and as complex as ethinyl estradiol (Section 2.1C) in birth control pills have this structural unit in common. All aldehydes have a —C(═O)— (carbonyl) group with one bond to a hydrogen and the other to one or more carbons, such as in benzaldehyde (from almonds). All ketones include a carbonyl group bonded by its carbon to one or more other carbons on each side, as in the natural oil menthone, found in geraniums and spearmint.



Ethanol



Benzaldehyde



Menthone

Members of each functional group family share common chemical properties and reactivity, and this fact helps greatly in organizing our knowledge of organic chemistry. As you progress in this chapter it will serve you well to learn the arrangements of atoms that define the common functional groups. This knowledge will be invaluable to your study of organic chemistry.

Toward the end of this chapter we introduce an instrumental technique called *infrared spectroscopy* that provides physical evidence for the presence of particular functional groups. You will very likely make use of infrared spectroscopy in your organic laboratory work.

Let us begin this chapter with families of compounds that contain only carbon and hydrogen.

2.1 *Hydrocarbons: Representative Alkanes, Alkenes, Alkynes, and Aromatic Compounds*

Here we introduce the class of compounds that contains only carbon and hydrogen, and we shall see how the -ane, -ene, or -yne ending in a name tells us what kinds of carbon–carbon bonds are present.

- **Hydrocarbons** are compounds that contain only carbon and hydrogen atoms.

Methane (CH_4) and ethane (C_2H_6) are hydrocarbons, for example. They also belong to a subgroup of compounds called alkanes.

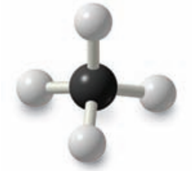
- **Alkanes** are hydrocarbons that do not have multiple bonds between carbon atoms, and we can indicate this in the family name and in names for specific compounds by the **-ane** ending.

Other hydrocarbons may contain double or triple bonds between their carbon atoms.

- **Alkenes** contain at least one carbon–carbon double bond, and this is indicated in the family name and in names for specific compounds by an **-ene** ending.
- **Alkynes** contain at least one carbon–carbon triple bond, and this is indicated in the family name and in names for specific compounds by a **-yne** ending.
- **Aromatic compounds** contain a special type of ring, the most common example of which is a benzene ring. There is no special ending for the general family of aromatic compounds.

We shall introduce representative examples of each of these classes of hydrocarbons in the following sections.

Generally speaking, compounds such as the alkanes, whose molecules contain only single bonds, are referred to as **saturated compounds** because these compounds contain the maximum number of hydrogen atoms that the carbon compound can possess. Compounds with multiple bonds, such as alkenes, alkynes, and aromatic hydrocarbons, are called **unsaturated compounds** because they possess fewer than the maximum number of hydrogen atoms, and they are capable of reacting with hydrogen under the proper conditions. We shall have more to say about this in Chapter 7.

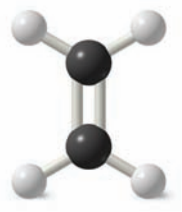

Methane

2.1A Alkanes

The primary sources of alkanes are natural gas and petroleum. The smaller alkanes (methane through butane) are gases under ambient conditions. Methane is the principal component of natural gas. Higher molecular weight alkanes are obtained largely by refining petroleum. Methane, the simplest alkane, was one major component of the early atmosphere of this planet. Methane is still found in Earth's atmosphere, but no longer in appreciable amounts. It is, however, a major component of the atmospheres of Jupiter, Saturn, Uranus, and Neptune.

Some living organisms produce methane from carbon dioxide and hydrogen. These very primitive creatures, called *methanogens*, may be Earth's oldest organisms, and they may represent a separate form of evolutionary development. Methanogens can survive only in an anaerobic (i.e., oxygen-free) environment. They have been found in ocean trenches, in mud, in sewage, and in cows' stomachs.

2.1B Alkenes



Ethene

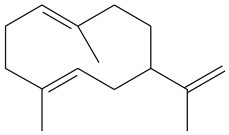
Ethene and propene, the two simplest alkenes, are among the most important industrial chemicals produced in the United States. Each year, the chemical industry produces more than 30 billion pounds of ethene and about 15 billion pounds of propene. Ethene is used as a starting material for the synthesis of many industrial compounds, including ethanol, ethylene oxide, ethanal, and the polymer polyethylene (Section 10.10). Propene is used in making the polymer polypropylene (Section 10.10 and Special Topic B*), and, in addition to other uses, propene is the starting material for a synthesis of acetone and cumene (Section 21.4B).

Ethene also occurs in nature as a plant hormone. It is produced naturally by fruits such as tomatoes and bananas and is involved in the ripening process of these fruits. Much use is now made of ethene in the commercial fruit industry to bring about the ripening of tomatoes and bananas picked green because the green fruits are less susceptible to damage during shipping.

Ethene ripens bananas.

There are many naturally occurring alkenes. Two examples are the following:

β-Pinene (a component of turpentine)



An aphid alarm pheromone

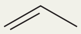
Solved Problem 2.1

Propene, $CH_3CH{=}CH_2$, is an alkene. Write the structure of a constitutional isomer of propene that is not an alkene. (*Hint*: It does not have a double bond.)

STRATEGY AND ANSWER A compound with a ring of *n* carbon atoms will have the same molecular formula as an alkene with the same number of carbons.

Cyclopropane
C_3H_6

is a constitutional isomer of



Propene
C_3H_6

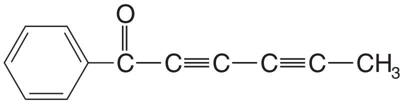
Cyclopropane has anesthetic properties.

2.1C Alkynes

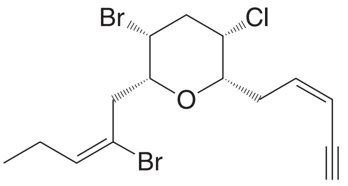
Ethyne

The simplest alkyne is ethyne (also called acetylene). Alkynes occur in nature and can be synthesized in the laboratory.

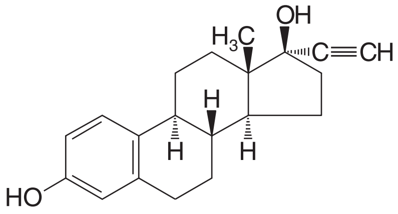
Two examples of alkynes among thousands that have a biosynthetic origin are capillin, an antifungal agent, and dactylyne, a marine natural product that is an inhibitor of pentobarbital metabolism. Ethinyl estradiol is a synthetic alkyne whose estrogen-like properties have found use in oral contraceptives.



Capillin



Dactylyne

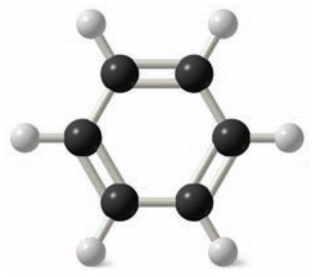


Ethinyl estradiol
[17α-ethynyl-1,3,5(10)-estratriene-3,17β-diol]

---

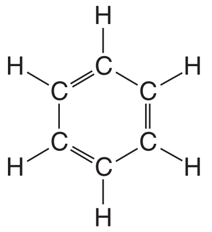
*Special Topics A-F and H are in WileyPLUS; Special Topic G can be found later in this volume.

2.1D Benzene: A Representative Aromatic Hydrocarbon



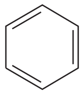
Benzene

In Chapter 14 we shall study in detail a group of unsaturated cyclic hydrocarbons known as aromatic compounds. The compound known as **benzene** is the prototypical aromatic compound. Benzene can be written as a six-membered ring with alternating single and double bonds, called a **Kekulé structure** after August Kekulé (Section 1.3), who first conceived of this representation:



Kekulé structure for benzene

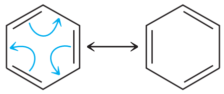
or



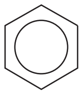
Bond-line representation of Kekulé structure

Even though the Kekulé structure is frequently used for benzene compounds, there is much evidence that this representation is inadequate and incorrect. For example, if benzene had alternating single and double bonds as the Kekulé structure indicates, we would expect the lengths of the carbon–carbon bonds around the ring to be alternately longer and shorter, as we typically find with carbon–carbon single and double bonds (Fig. 1.31). In fact, the carbon–carbon bonds of benzene are all the same length (1.39 Å), a value in between that of a carbon–carbon single bond and a carbon–carbon double bond. There are two ways of dealing with this problem: with resonance theory or with molecular orbital theory.

If we use resonance theory, we visualize benzene as being represented by either of two equivalent Kekulé structures:



Two contributing Kekulé structures for benzene

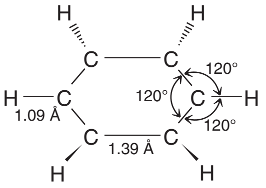


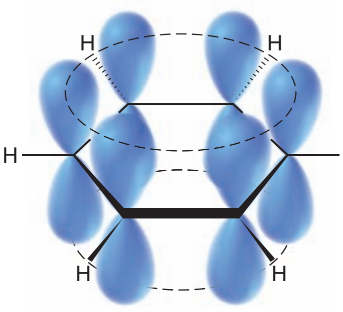
A representation of the resonance hybrid

Based on the principles of resonance theory (Section 1.8) we recognize that benzene cannot be represented adequately by either structure, but that, instead, *it should be visualized as a hybrid of the two structures*. We represent this hybrid by a hexagon with a circle in the middle. Resonance theory, therefore, solves the problem we encountered in understanding how all of the carbon–carbon bonds are the same length. According to resonance theory, the bonds are not alternating single and double bonds, they are a resonance hybrid of the two: Any bond that is a single bond in the first contributor is a double bond in the second, and vice versa.

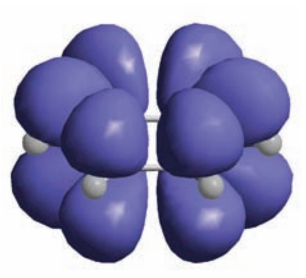
- All of the carbon–carbon bonds in benzene are one and one-half bonds, have a bond length in between that of a single bond and a double bond, and have bond angles of 120°.

In the molecular orbital explanation, which we shall describe in much more depth in Chapter 14, we begin by recognizing that the carbon atoms of the benzene ring are sp^2 hybridized. Therefore, each carbon has a p orbital that has one lobe above the plane of the ring and one lobe below, as shown here in the schematic and calculated p orbital representations.





Schematic representation of benzene *p* orbitals



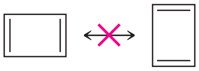
Calculated *p* orbital shapes in benzene

Calculated benzene molecular orbital resulting from favorable overlap of *p* orbitals above and below plane of benzene ring

The lobes of each *p* orbital above and below the ring overlap with the lobes of *p* orbitals on the atoms to either side of it. This kind of overlap of *p* orbitals leads to a set of bonding molecular orbitals that encompass all of the carbon atoms of the ring, as shown in the calculated molecular orbital. Therefore, the six electrons associated with these *p* orbitals (one electron from each orbital) are **delocalized** about all six carbon atoms of the ring. This delocalization of electrons explains how all the carbon–carbon bonds are equivalent and have the same length. In Section 14.7B, when we study nuclear magnetic resonance spectroscopy, we shall present convincing physical evidence for this delocalization of the electrons.

Review Problem 2.1

Cyclobutadiene (below) is like benzene in that it has alternating single and double bonds in a ring. However, its bonds are not the same length, the double bonds being shorter than the single bonds; the molecule is rectangular, not square. Explain why it would be incorrect to write resonance structures as shown.

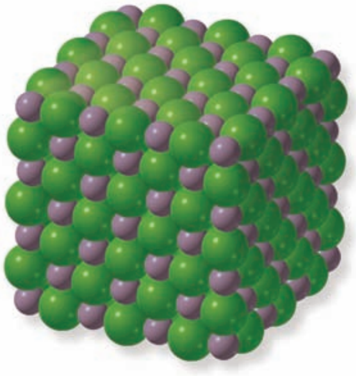


2.2 *Polar Covalent Bonds*

In our discussion of chemical bonds in Section 1.4, we examined compounds such as LiF in which the bond is between two atoms with very large electronegativity differences. In instances like these, we said, a complete transfer of electrons occurs, giving the compound an **ionic bond**:

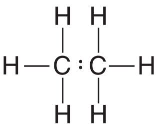
$$\mathrm{Li}^{+}\ :\!\ddot{\underset{\cdot\cdot}{\mathrm{F}}}\!:^{-}$$

Lithium fluoride has an ionic bond.



Lithium fluoride crystal model.

We also described molecules in which electronegativity differences are not large, or in which they are the same, such as the carbon–carbon bond of ethane. Here the electrons are shared equally between the atoms.



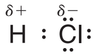
$$\mathrm{H{-}\underset{\underset{\displaystyle H}{|}}{\overset{\overset{\displaystyle H}{|}}{C}}\!:\!\underset{\underset{\displaystyle H}{|}}{\overset{\overset{\displaystyle H}{|}}{C}}{-}H}$$

Ethane has a covalent bond. The electrons are shared equally between the carbon atoms.

Until now, we have not considered the possibility that the electrons of a covalent bond might be shared unequally.

- If electronegativity differences exist between two bonded atoms, and they are not large, the electrons are not shared equally and a **polar covalent bond** is the result.
- Remember: One definition of **electronegativity** is *the ability of an atom to attract electrons that it is sharing in a covalent bond.*

An example of such a polar covalent bond is the one in hydrogen chloride. The chlorine atom, with its greater electronegativity, pulls the bonding electrons closer to it. This makes the hydrogen atom somewhat electron deficient and gives it a *partial* positive charge ($\delta+$). The chlorine atom becomes somewhat electron rich and bears a *partial* negative charge ($\delta-$):

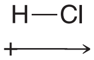


Because the hydrogen chloride molecule has a partially positive end and a partially negative end, it is a **dipole**, and it has a **dipole moment**.

The direction of polarity of a polar bond can be symbolized by a vector quantity +⟶. The crossed end of the arrow is the positive end and the arrowhead is the negative end:

(positive end) +⟶ (negative end)

In HCl, for example, we would indicate the direction of the dipole moment in the following way:



The dipole moment is a physical property that can be measured experimentally. It is defined as the product of the magnitude of the charge in electrostatic units (esu) and the distance that separates them in centimeters (cm):

$$\text{Dipole moment} = \text{charge (in esu)} \times \text{distance (in cm)}$$
$$\mu = e \times d$$

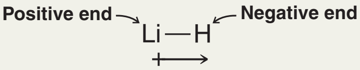
The charges are typically on the order of 10^{-10} esu and the distances are on the order of 10^{-8} cm. Dipole moments, therefore, are typically on the order of 10^{-18} esu·cm. For convenience, this unit, 1×10^{-18} esu cm, is defined as one **debye** and is abbreviated D. (The unit is named after Peter J. W. Debye, a chemist born in the Netherlands, who taught at Cornell University from 1936 to 1966. Debye won the Nobel Prize in Chemistry in 1936.) In SI units $1\ \text{D} = 3.336 \times 10^{-30}$ coulomb meter (C · m).

If necessary, the length of the arrow can be used to indicate the magnitude of the dipole moment. Dipole moments, as we shall see in Section 2.3, are very useful quantities in accounting for physical properties of compounds.

Solved Problem 2.2

Using a dipole moment arrow as shown above and the table of electronegativities (Table 1.2), indicate the direction of the dipole moment of lithium hydride (LiH), a covalent compound. Also place $\delta+$ and $\delta-$ symbols near the Li and H as appropriate.

STRATEGY AND ANSWER In Table 1.2 we find that lithium (a metal) has a very low electronegativity of 1.0. Hydrogen (a nonmetal) has a larger electronegativity of 2.1. The hydrogen atom will pull the electrons it is sharing with lithium in its direction. The bond between lithium and hydrogen will be polar with the lithium at the positive end and the hydrogen at the negative end.



Using $\delta+$ and $\delta-$ symbols we can show the polarity of LiF as follows:

$$^{\delta+}\text{Li—F}^{\delta-}$$

Review Problem 2.2

Write $\delta+$ and $\delta-$ by the appropriate atoms and draw a dipole moment vector for any of the following molecules that are polar:

(a) HF **(b)** IBr **(c)** Br_2 **(d)** F_2

Polar covalent bonds strongly influence the physical properties and reactivity of molecules. In many cases, these polar covalent bonds are part of **functional groups**, which we shall study shortly (Sections 2.5–2.13). Functional groups are defined groups of atoms in a molecule that give rise to the function (reactivity or physical properties) of the molecule. Functional groups often contain atoms having different electronegativity values and unshared electron pairs. (Atoms such as oxygen, nitrogen, and sulfur that form covalent bonds and have unshared electron pairs are called **heteroatoms**.)

2.2A Maps of Electrostatic Potential

One way to visualize the distribution of charge in a molecule is with a **map of electrostatic potential (MEP)**. Regions of an electron density surface that are more negative than others in an MEP are colored red. These regions would attract a positively charged species (or repel a negative charge). Regions in the MEP that are less negative (or are positive) are blue. Blue regions are likely to attract electrons from another molecule. The spectrum of colors from red to blue indicates the trend in charge from most negative to least negative (or most positive).

Figure 2.1 shows a map of electrostatic potential for the low-electron-density surface of hydrogen chloride. We can see clearly that negative charge is concentrated near the chlorine atom and that positive charge is localized near the hydrogen atom, as we predict based on the difference in their electronegativity values. Furthermore, because this MEP is plotted at the low-electron-density surface of the molecule (the van der Waals surface, Section 2.13B), it also gives an indication of the molecule's overall shape.

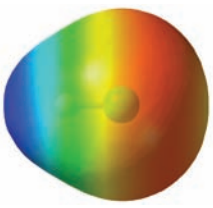


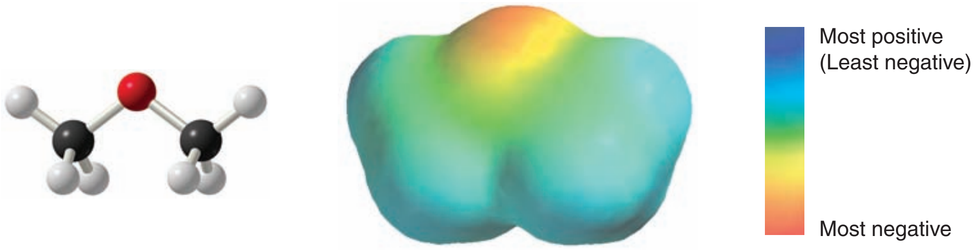
Figure 2.1 A calculated map of electrostatic potential for hydrogen chloride showing regions of relatively more negative charge in red and more positive charge in blue. Negative charge is clearly localized near the chlorine, resulting in a strong dipole moment for the molecule.

THE CHEMISTRY OF . . .

Calculated Molecular Models: Maps of Electrostatic Potential

A map of electrostatic potential is prepared by carrying out a quantum mechanical calculation that involves moving an imaginary positive point charge at a fixed distance over a given electron density surface of a molecule. As this is done, the varying potential energy in the attraction between the electron cloud and the imaginary positive charge is plotted in color-coded fashion. Red in the MEP indicates strong attraction between the electron density surface at that location and the probing positive charge—in other words, greater negative charge at that part of the surface. Blue regions in the map indicate weaker attraction between the surface and the positive charge probe. The overall distribution of charge is indicated by the trend from blue (most positive or least negative) to green or yellow (neutral) to red (most negative). Most often we plot the MEP at the van der Waals surface of a molecule since that represents approximately the furthest extent of a molecule's electron cloud and therefore its overall shape. The molecular model in this box is for dimethyl ether. The MEP shows the concentration of negative charge where the unshared electron pairs are located on the oxygen atom.

It is important to note that when directly comparing the MEP for one molecule to that of another, the color scheme used to represent the charge scale in each model must be the same. When we make direct comparisons between molecules, we will plot their MEPs on the same scale. We will find that such comparisons are especially useful because they allow us to compare the electron distribution in one molecule to that in another and predict how one molecule might interact with the electrons of another molecule.



Dimethyl ether

Helpful Hint

Electron density surfaces and electrostatic potential maps.

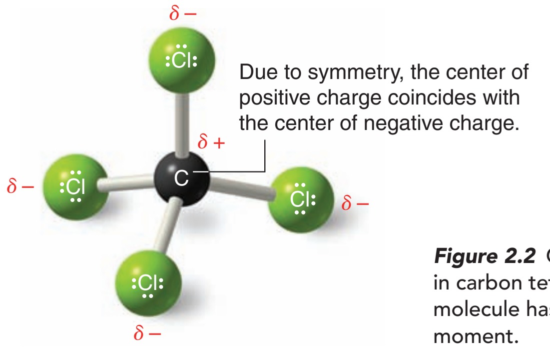
2.3 *Polar and Nonpolar Molecules*

In the discussion of dipole moments in the previous section, our attention was restricted to simple diatomic molecules. Any *diatomic* molecule in which the two atoms are *different* (and thus have different electronegativities) will, of necessity, have a dipole moment. In general, a molecule with a dipole moment is a **polar molecule**. If we examine Table 2.1, however, we find that a number of molecules (e.g., CCl_4, CO_2) consist of more than two atoms, have *polar* bonds, *but have no dipole moment*. With our knowledge of the shapes of molecules (Sections 1.12–1.16) we can understand how this can occur.

TABLE 2.1 Dipole Moments of Some Simple Molecules

Formula	μ (D)	Formula	μ (D)
H_2	0	CH_4	0
Cl_2	0	CH_3Cl	1.87
HF	1.83	CH_2Cl_2	1.55
HCl	1.08	$CHCl_3$	1.02
HBr	0.80	CCl_4	0
HI	0.42	NH_3	1.47
BF_3	0	NF_3	0.24
CO_2	0	H_2O	1.85

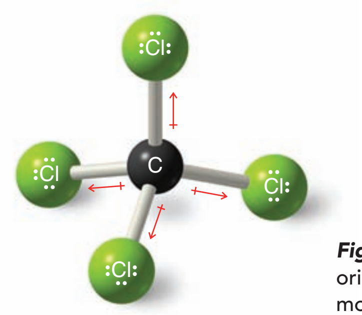
Consider a molecule of carbon tetrachloride (CCl_4). Because the electronegativity of chlorine is greater than that of carbon, each of the carbon–chlorine bonds in CCl_4 is polar. Each chlorine atom has a partial negative charge, and the carbon atom is considerably positive. Because a molecule of carbon tetrachloride is tetrahedral (Fig. 2.2), however, *the center of positive charge and the center of negative charge coincide, and the molecule has no net dipole moment*.



Due to symmetry, the center of positive charge coincides with the center of negative charge.

Figure 2.2 Charge distribution in carbon tetrachloride. The molecule has no net dipole moment.

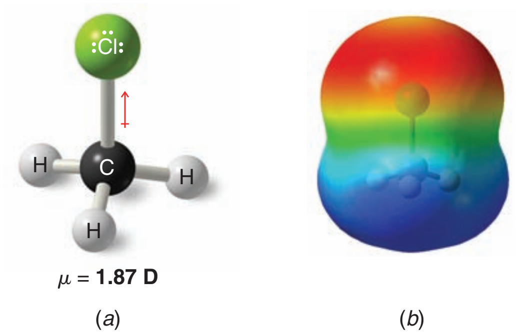
This result can be illustrated in a slightly different way: If we use arrows ($+\!\!\longrightarrow$) to represent the direction of polarity of each bond, we get the arrangement of bond moments shown in Fig. 2.3. Since the bond moments are vectors of equal magnitude arranged tetrahedrally, their effects cancel. Their vector sum is zero. The molecule has *no net dipole moment*.



μ = **0 D**

Figure 2.3 A tetrahedral orientation of equal bond moments causes their effects to cancel.

The chloromethane molecule (CH_3Cl) has a net dipole moment of 1.87 D. Since carbon and hydrogen have electronegativities (Table 1.2) that are nearly the same, the contribution of three C—H bonds to the net dipole is negligible. The electronegativity difference between carbon and chlorine is large, however, and the highly polar C—Cl bond accounts for most of the dipole moment of CH_3Cl (Fig. 2.4).



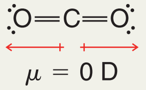
(a) (b)

Figure 2.4 *(a)* The dipole moment of chloromethane arises mainly from the highly polar carbon–chlorine bond. *(b)* A map of electrostatic potential illustrates the polarity of chloromethane.

Solved Problem 2.3

Although molecules of CO_2 have polar bonds (oxygen is more electronegative than carbon), carbon dioxide (Table 2.1) has no dipole moment. What can you conclude about the geometry of a carbon dioxide molecule?

STRATEGY AND ANSWER For a CO_2 molecule to have a zero dipole moment, the bond moments of the two carbon–oxygen bonds must cancel each other. This can happen only if molecules of carbon dioxide are linear.



Review Problem 2.3

Boron trifluoride (BF_3) has no dipole moment ($\mu = 0$ D). Explain how this observation confirms the geometry of BF_3 predicted by VSEPR theory.

Review Problem 2.4

Tetrachloroethene ($CCl_2{=}CCl_2$) does not have a dipole moment. Explain this fact on the basis of the shape of $CCl_2{=}CCl_2$.

Review Problem 2.5

Sulfur dioxide (SO_2) has a dipole moment ($\mu = 1.63$ D); on the other hand, carbon dioxide (see Solved Problem 2.3) has no dipole moment ($\mu = 0$ D). What do these facts indicate about the geometry of sulfur dioxide?

Unshared pairs of electrons make large contributions to the dipole moments of water and ammonia. Because an unshared pair has no other atom attached to it to partially neutralize its negative charge, an unshared electron pair contributes a large moment directed away from the central atom (Fig. 2.5). (The O—H and N—H moments are also appreciable.)

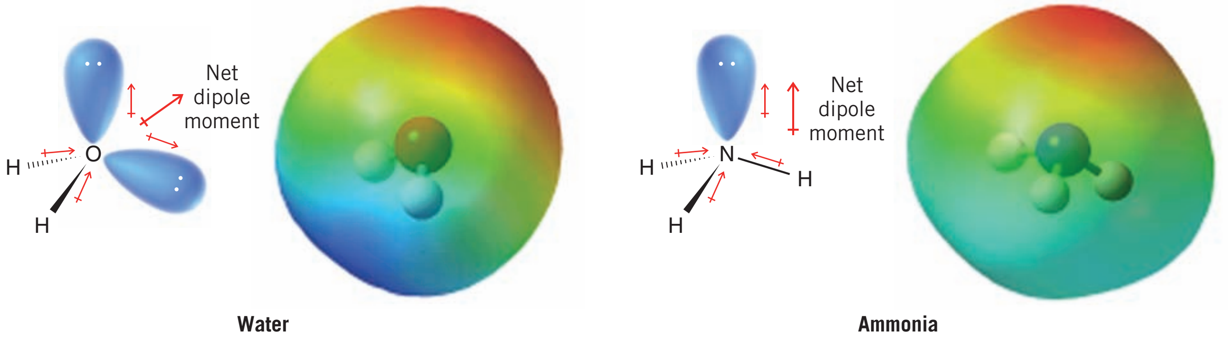


Figure 2.5 Bond moments and the resulting dipole moments of water and ammonia.

Review Problem 2.6

Using a three-dimensional formula, show the direction of the dipole moment of CH_3OH. Write $\delta+$ and $\delta-$ signs next to the appropriate atoms.

Review Problem 2.7

Trichloromethane ($CHCl_3$, also called *chloroform*) has a larger dipole moment than $CFCl_3$. Use three-dimensional structures and bond moments to explain this fact.

2.3A Dipole Moments in Alkenes

Cis–trans isomers of alkenes (Section 1.13B) have different physical properties. They have different melting points and boiling points, and often cis–trans isomers differ markedly in the magnitude of their dipole moments. Table 2.2 summarizes some of the physical properties of two pairs of cis–trans isomers.

TABLE 2.2 Physical Properties of Some Cis–Trans Isomers

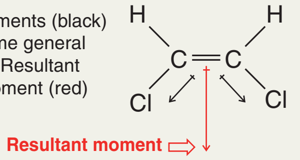
Compound	Melting Point (°C)	Boiling Point (°C)	Dipole Moment (D)
cis-1,2-Dichloroethene	−80	60	1.90
trans-1,2-Dichloroethene	−50	48	0
cis-1,2-Dibromoethene	−53	112.5	1.35
trans-1,2-Dibromoethene	−6	108	0

Solved Problem 2.4

Explain why *cis*-1,2-dichloroethene (Table 2.2) has a large dipole moment whereas *trans*-1,2-dichloroethene has a dipole moment equal to zero.

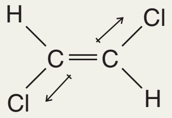
STRATEGY AND ANSWER If we examine the net dipole moments (shown in red) for the bond moments (black), we see that in *trans*-1,2-dichloroethene the bond moments cancel each other, whereas in *cis*-1,2-dichloroethene they augment each other.

Bond moments (black) are in same general direction. Resultant dipole moment (red) is large.



***cis*-1,2-Dichloroethene**

μ = 1.9 D



Bond moments cancel each other. Net dipole is zero.

***trans*-1,2-Dichloroethene**

μ = 0 D

Review Problem 2.8

Indicate the direction of the important bond moments in each of the following compounds (neglect C—H bonds). You should also give the direction of the net dipole moment for the molecule. If there is no net dipole moment, state that $\mu = 0$ D.

(a) *cis*-CHF═CHF **(b)** *trans*-CHF═CHF **(c)** $CH_2{=}CF_2$ **(d)** $CF_2{=}CF_2$

Review Problem 2.9

Write structural formulas for all of the alkenes with **(a)** the formula $C_2H_2Br_2$ and **(b)** the formula $C_2Br_2Cl_2$. In each instance designate compounds that are cis–trans isomers of each other. Predict the dipole moment of each one.

2.4 *Functional Groups*

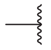
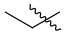
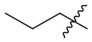
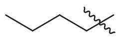
- **Functional groups** are common and specific arrangements of atoms that impart predictable reactivity and properties to a molecule.

The functional group of an alkene, for example, is its carbon–carbon double bond. When we study the reactions of alkenes in greater detail in Chapter 8, we shall find that most of the chemical reactions of alkenes are the chemical reactions of the carbon–carbon double bond.

The functional group of an alkyne is its carbon–carbon triple bond. Alkanes do not have a functional group. Their molecules have carbon–carbon single bonds and carbon–hydrogen bonds, but these bonds are present in molecules of almost all organic compounds, and C—C and C—H bonds are, in general, much less reactive than common functional groups. We shall introduce other common functional groups and their properties in Sections 2.5–2.11. Table 2.3 (Section 2.12) summarizes the most important functional groups. First, however, let us introduce some common alkyl groups, which are specific groups of carbon and hydrogen atoms that are not part of functional groups.

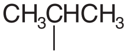
2.4A Alkyl Groups and the Symbol R

Alkyl groups are the groups that we identify for purposes of naming compounds. They are groups that would be obtained by removing a hydrogen atom from an alkane:

ALKANE	ALKYL GROUP	ABBREVIATION	BOND-LINE	MODEL
$CH_3—H$ **Methane**	$CH_3—$ **Methyl**	Me-		
$CH_3CH_2—H$ **Ethane**	$CH_3CH_2—$ **Ethyl**	Et-		
$CH_3CH_2CH_2—H$ **Propane**	$CH_3CH_2CH_2—$ **Propyl**	Pr-		
$CH_3CH_2CH_2CH_2—H$ **Butane**	$CH_3CH_2CH_2CH_2—$ **Butyl**	Bu-		

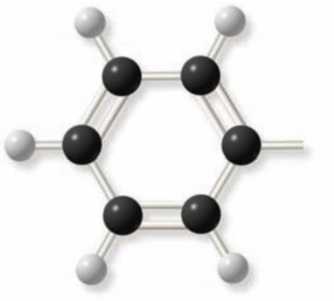
While only one alkyl group can be derived from methane or ethane (the **methyl** and **ethyl** groups, respectively), two groups can be derived from propane. Removal of a hydrogen from one of the end carbon atoms gives a group that is called the **propyl** group; removal of a hydrogen from the middle carbon atom gives a group that is called the **isopropyl** group. The names and structures of these groups are used so frequently in organic chemistry that you should learn them now. See Section 4.3C for names and structures of branched alkyl groups derived from butane and other hydrocarbons.

We can simplify much of our future discussion if, at this point, we introduce a symbol that is widely used in designating general structures of organic molecules: the symbol R. R ***is used as a general symbol to represent any alkyl group.*** For example, R might be a methyl group, an ethyl group, a propyl group, or an isopropyl group:

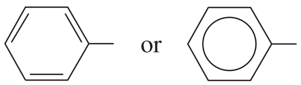
$CH_3—$	Methyl	**These and others can be designated by R.**
$CH_2CH_2—$	Ethyl	
$CH_3CH_2CH_2—$	Propyl	
 $CH_3\underset{|}{C}HCH_3$	Isopropyl	

Thus, the general formula for an alkane is R—H.

2.4B Phenyl and Benzyl Groups



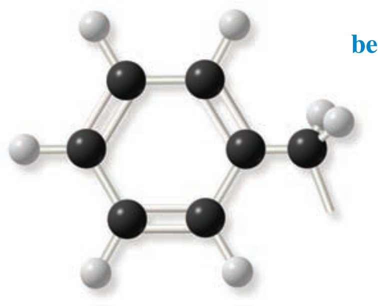
When a benzene ring is attached to some other group of atoms in a molecule, it is called a **phenyl group**, and it is represented in several ways:



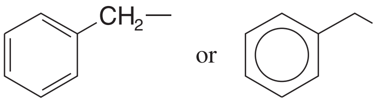
or C_6H_5— or Ph— or

ϕ— or Ar— (if ring substituents are present)

Ways of representing a phenyl group



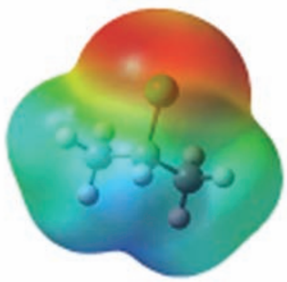
The combination of a phenyl group and a **methylene group** (—CH_2—) is called a **benzyl group**:



or $C_6H_5CH_2$— or Bn—

Ways of representing a benzyl group

2.5 *Alkyl Halides or Haloalkanes*



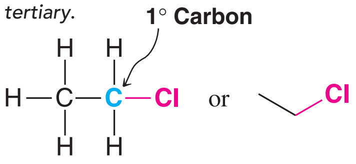
2-Chloropropane

Alkyl halides are compounds in which a halogen atom (fluorine, chlorine, bromine, or iodine) replaces a hydrogen atom of an alkane. For example, CH_3Cl and CH_3CH_2Br are alkyl halides. Alkyl halides are also called **haloalkanes**. The generic formula for an alkyl halide is $R-\ddot{\underset{..}{X}}:$ where X = fluorine, chlorine, bromine, or iodine.

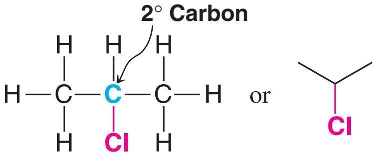
Alkyl halides are classified as being primary (1°), secondary (2°), or tertiary (3°). ***This classification is based on the carbon atom to which the halogen is directly attached***. If the carbon *atom* that bears the halogen is attached to only one other carbon, the carbon atom is said to be a **primary carbon atom** and the alkyl halide is classified as a **primary alkyl halide**. If the carbon that bears the halogen is itself attached to two other carbon atoms, then the carbon is a **secondary carbon** and the alkyl halide is a **secondary alkyl halide**. If the carbon that bears the halogen is attached to three other carbon atoms, then the carbon is a **tertiary carbon** and the alkyl halide is a **tertiary alkyl halide**. Examples of primary, secondary, and tertiary alkyl halides are the following:

Helpful Hint

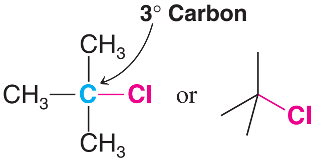
Although we use the symbols 1°, 2°, 3°, we do not *say* first degree, second degree, and third degree; we say *primary*, *secondary*, and *tertiary*.



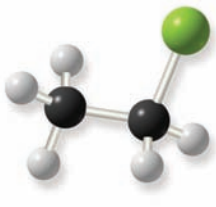
A 1° alkyl chloride

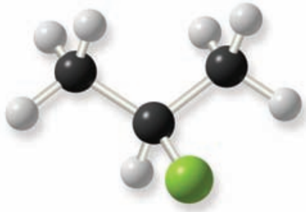


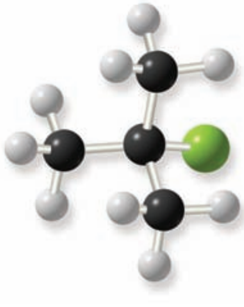
A 2° alkyl chloride



A 3° alkyl chloride



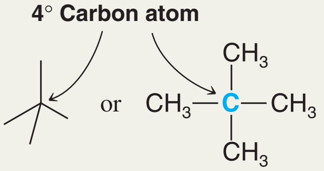




Solved Problem 2.5

Write the structure of an alkane with the formula C_5H_{12} that has no secondary or tertiary carbon atoms. *Hint*: The compound has a quaternary (4°) carbon.

STRATEGY AND ANSWER Following the pattern of designations for carbon atoms given above, a 4° carbon atom must be one that is attached to four other carbon atoms. If we start with this carbon atom, and then add four carbon atoms with their attached hydrogens, there is only one possible alkane. The other four carbons are all primary carbons; none is secondary or tertiary.



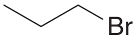
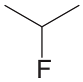
$$CH_3-\overset{\overset{\large CH_3}{|}}{\underset{\underset{\large CH_3}{|}}{C}}-CH_3$$

Review Problem 2.10

Write bond-line structural formulas for **(a)** two constitutionally isomeric primary alkyl bromides with the formula C_4H_9Br, **(b)** a secondary alkyl bromide, and **(c)** a tertiary alkyl bromide with the same formula. Build handheld molecular models for each structure and examine the differences in their connectivity.

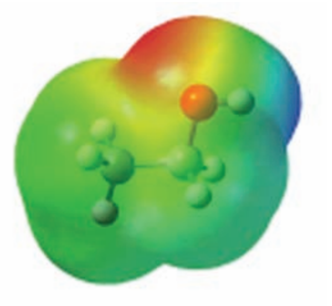
Review Problem 2.11

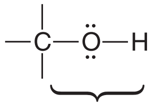
Although we shall discuss the naming of organic compounds later when we discuss the individual families in detail, one method of naming alkyl halides is so straightforward that it is worth describing here. We simply name the alkyl group attached to the halogen and add the word *fluoride, chloride, bromide*, or *iodide*. Write formulas for **(a)** ethyl fluoride and **(b)** isopropyl chloride.

What are the names for **(c)** , **(d)** , and **(e)** C_6H_5I?

2.6 *Alcohols*

Methyl alcohol (also called methanol) has the structural formula CH_3OH and is the simplest member of a family of organic compounds known as **alcohols**. The characteristic functional group of this family is the hydroxyl (—OH) group attached to an sp^3-hybridized carbon atom. Another example of an alcohol is ethyl alcohol, CH_3CH_2OH (also called ethanol).

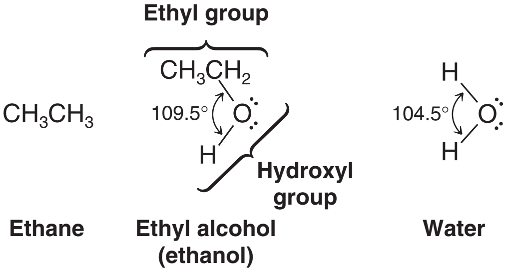

Ethanol



$$-\overset{|}{\underset{|}{C}}-\ddot{\underset{\cdot\cdot}{O}}-H$$

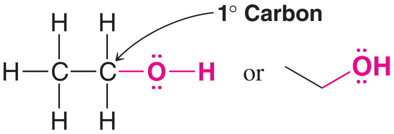
This is the functional group of an alcohol.

Alcohols may be viewed structurally in two ways: (1) as hydroxyl derivatives of alkanes and (2) as alkyl derivatives of water. Ethyl alcohol, for example, can be seen as an ethane molecule in which one hydrogen has been replaced by a hydroxyl group or as a water molecule in which one hydrogen has been replaced by an ethyl group:

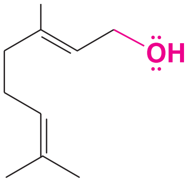


As with alkyl halides, alcohols are classified into three groups: primary (1°), secondary (2°), and tertiary (3°) alcohols. ***This classification is based on the degree of substitution of the carbon to which the hydroxyl group is directly attached.*** If the carbon has only one other carbon attached to it, the carbon is said to be a **primary carbon** and the alcohol is a **primary alcohol**:

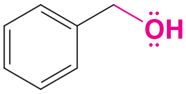

Geraniol is a major component of the oil of roses.



Ethyl alcohol (a 1° alcohol)

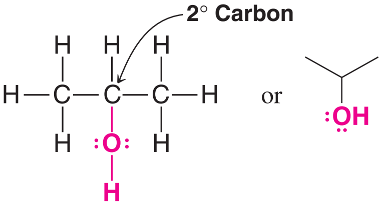


Geraniol (a 1° alcohol)

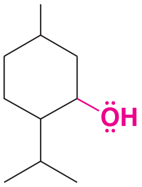


Benzyl alcohol (a 1° alcohol)

If the carbon atom that bears the hydroxyl group also has two other carbon atoms attached to it, this carbon is called a secondary carbon, and the alcohol is a secondary alcohol:



Isopropyl alcohol (a 2° alcohol)

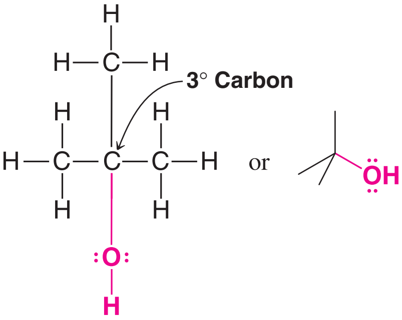


Menthol (a 2° alcohol found in peppermint oil)

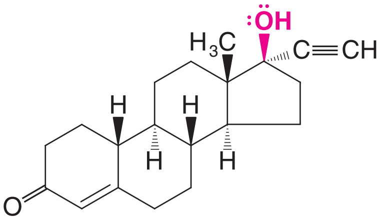
If the carbon atom that bears the hydroxyl group has three other carbons attached to it, this carbon is called a tertiary carbon, and the alcohol is a tertiary alcohol:

Helpful Hint

Practice with handheld molecular models by building models of as many of the compounds on this page as you can.



***tert*-Butyl alcohol (a 3° alcohol)**



Norethindrone (an oral contraceptive that contains a 3° alcohol group as well as a ketone group and carbon–carbon double and triple bonds)

Norethindrone birth control pills.

Review Problem 2.12

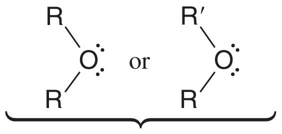
Write bond-line structural formulas for **(a)** two primary alcohols, **(b)** a secondary alcohol, and **(c)** a tertiary alcohol—all having the molecular formula $C_4H_{10}O$.

Review Problem 2.13

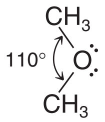
One way of naming alcohols is to name the alkyl group that is attached to the —OH and add the word *alcohol*. Write bond-line formulas for **(a)** propyl alcohol and **(b)** isopropyl alcohol.

2.7 Ethers

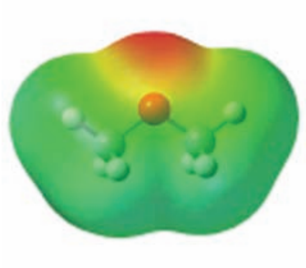
Ethers have the general formula R—O—R or R—O—R′, where R′ may be an alkyl (or phenyl) group different from R. Ethers can be thought of as derivatives of water in which both hydrogen atoms have been replaced by alkyl groups. The bond angle at the oxygen atom of an ether is only slightly larger than that of water:



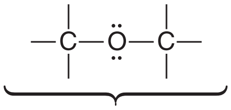
General formula for an ether



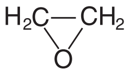
Dimethyl ether (a typical ether)



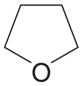
Dimethyl ether



The functional group of an ether



Ethylene oxide

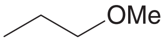


Tetrahydrofuran (THF)

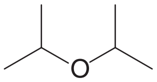
Two cyclic ethers

Review Problem 2.14

One way of naming ethers is to name the two alkyl groups attached to the oxygen atom in alphabetical order and add the word *ether*. If the two alkyl groups are the same, we use the prefix *di-*, for example, as in *dimethyl ether*. Write bond-line structural formulas for **(a)** diethyl ether, **(b)** ethyl propyl ether, and **(c)** ethyl isopropyl ether. What name would you give to **(d)**



(e)



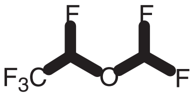
and **(f)** $CH_3OC_6H_5$?

THE CHEMISTRY OF . . .

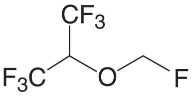
Ethers as General Anesthetics

Nitrous oxide (N_2O), also called laughing gas, was first used as an anesthetic in 1799, and it is still in use today, even though when used alone it does not produce deep anesthesia. The first use of an ether, diethyl ether, to produce deep anesthesia occurred in 1842. In the years that have passed since then, several different ethers, usually with halogen substituents, have replaced diethyl ether as anesthetics of choice. One reason: unlike diethyl ether, which is highly flammable, the halogenated ethers are not. Two halogenated ethers that are currently used for inhalation anesthesia are sevoflurane and desflurane.

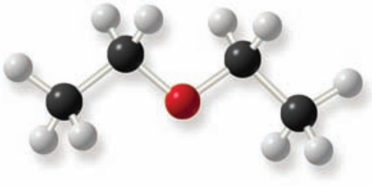
Diethyl ether

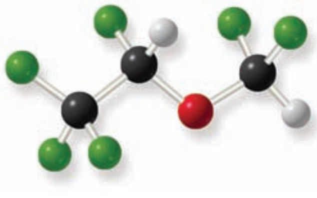


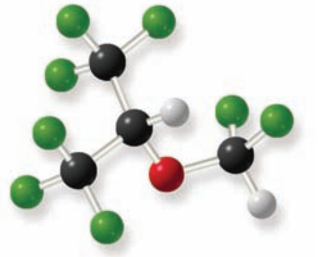
Desflurane



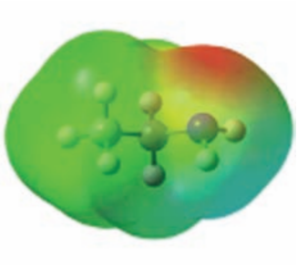
Sevoflurane





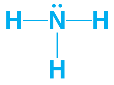


2.8 Amines



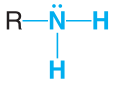
Ethylamine

Just as alcohols and ethers may be considered as organic derivatives of water, amines may be considered as organic derivatives of ammonia:



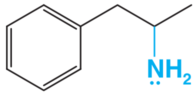
$$\begin{array}{c}\mathrm{H-\ddot{N}-H}\\ |\\ \mathrm{H}\end{array}$$

Ammonia

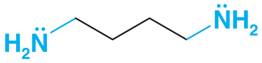


$$\begin{array}{c}\mathrm{R-\ddot{N}-H}\\ |\\ \mathrm{H}\end{array}$$

An amine

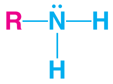


Amphetamine
(a dangerous stimulant)



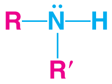
Putrescine
(found in decaying meat)

Amines are classified as primary, secondary, or tertiary amines. **This classification is based on *the number of organic groups that are attached to the nitrogen atom*:**



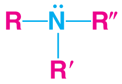
$$\begin{array}{c}\mathrm{R-\ddot{N}-H}\\ |\\ \mathrm{H}\end{array}$$

A primary (1°) amine



$$\begin{array}{c}\mathrm{R-\ddot{N}-H}\\ |\\ \mathrm{R'}\end{array}$$

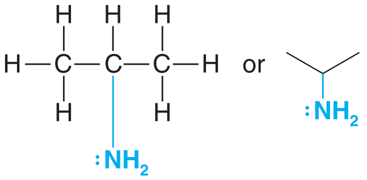
A secondary (2°) amine



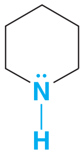
$$\begin{array}{c}\mathrm{R-\ddot{N}-R''}\\ |\\ \mathrm{R'}\end{array}$$

A tertiary (3°) amine

Notice that this is quite different from the way alcohols and alkyl halides are classified. Isopropylamine, for example, is a primary amine even though its —NH_2 group is attached to a secondary carbon atom. It is a primary amine because only one organic group is attached to the nitrogen atom:

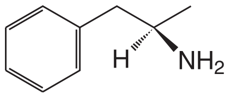


Isopropylamine
(a 1° amine)

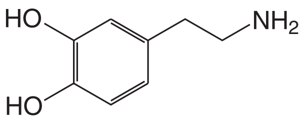


Piperidine
(a cyclic 2° amine)

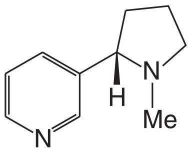
Amphetamine (below), a powerful and dangerous stimulant, is a primary amine. Dopamine, an important neurotransmitter whose depletion is associated with Parkinson's disease, is also a primary amine. Nicotine, a toxic compound found in tobacco that makes smoking addictive, has a secondary amine group and a tertiary one.



Amphetamine

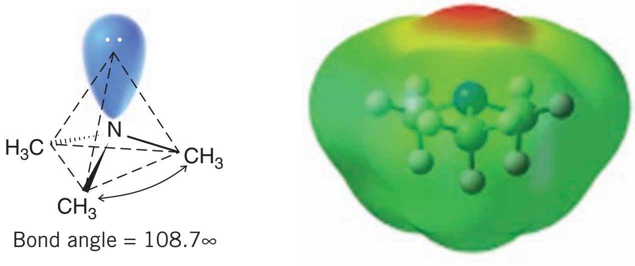


Dopamine



Nicotine

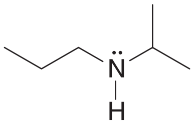
Amines are like ammonia (Section 1.16B) in having a trigonal pyramidal shape. The C—N—C bond angles of trimethylamine are 108.7°, a value very close to the H—C—H bond angles of methane. Thus, for all practical purposes, the nitrogen atom of an amine can be considered to be sp^3 hybridized with the unshared electron pair occupying one orbital (see below). This means that the unshared pair is relatively exposed, and as we shall see this is important because it is involved in almost all of the reactions of amines.

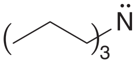


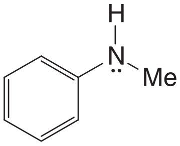
Trimethylamine

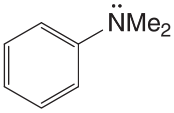
Review Problem 2.15

One way of naming amines is to name in alphabetical order the alkyl groups attached to the nitrogen atom, using the prefixes *di-* and *tri-* if the groups are the same. An example is *isopropylamine* whose formula is shown above. What are names for **(a)**, **(b)**, **(c)**, and **(d)**? Build handheld molecular models for the compounds in parts (a)–(d).

(a)


(b)


(c)


(d)


Write bond-line formulas for **(e)** propylamine, **(f)** trimethylamine, and **(g)** ethylisopropylmethylamine.

Review Problem 2.16

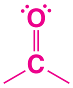
Which amines in Review Problem 2.15 are **(a)** primary amines, **(b)** secondary amines, and **(c)** tertiary amines?

Review Problem 2.17

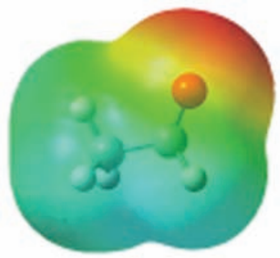
Amines are like ammonia in being weak bases. They do this by using their unshared electron pair to accept a proton. **(a)** Show the reaction that would take place between trimethylamine and HCl. **(b)** What hybridization state would you expect for the nitrogen atom in the product of this reaction?

2.9 Aldehydes and Ketones

Aldehydes and ketones both contain the **carbonyl group**—a group in which a carbon atom has a double bond to oxygen:



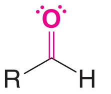
The carbonyl group



Acetaldehyde

The carbonyl group of an aldehyde is bonded to one hydrogen atom and one carbon atom (except for formaldehyde, which is the only aldehyde bearing two hydrogen atoms). The carbonyl group of a ketone is bonded to two carbon atoms. Using R, we can designate the general formulas for aldehydes and ketones as follows:

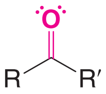
ALDEHYDES



or RCHO

(R = H in formaldehyde)

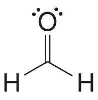
KETONES



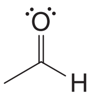
or RCOR′

(where R′ is an alkyl group that may be the same or different from R)

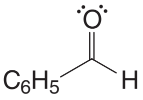
Some specific examples of aldehydes and ketones are the following:



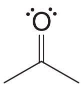
Formaldehyde



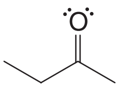
Acetaldehyde



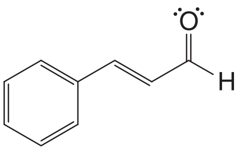
Benzaldehyde



Acetone

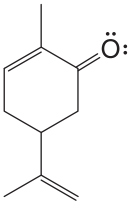


Ethyl methyl ketone



***trans*-Cinnamaldehyde (present in cinnamon)**

Cinnamon



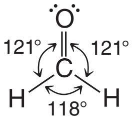
Carvone (from spearmint)

Spearmint

Helpful Hint

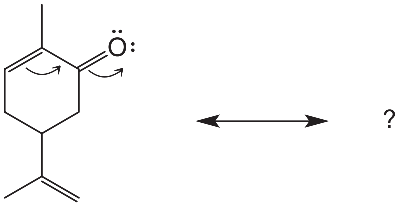
Computer molecular models can be found in the 3D Models section of the book's website for these and many other compounds we discuss in this book.

Aldehydes and ketones have a trigonal planar arrangement of groups around the carbonyl carbon atom. The carbon atom is *sp*^{2} hybridized. In formaldehyde, for example, the bond angles are as follows:

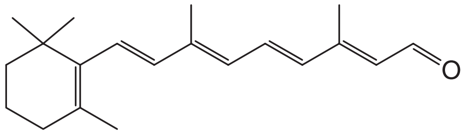


Review Problem 2.18

Write the resonance structure for carvone that results from moving the electrons as indicated. Include all formal charges.



Retinal (below) is an aldehyde made from vitamin A that plays a vital role in vision. We discuss this further in Chapter 13.



Retinal

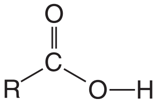
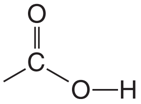
Review Problem 2.19

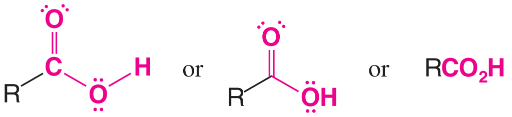
Write bond-line formulas for (**a**) four aldehydes and (**b**) three ketones that have the formula $C_5H_{10}O$.

2.10 *Carboxylic Acids, Esters, and Amides*

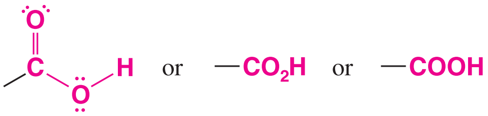
Carboxylic acids, esters, and amides all contain a carbonyl group that is bonded to an oxygen or nitrogen atom. As we shall learn in later chapters, all of these functional groups are interconvertible by appropriately chosen reactions.

2.10A Carboxylic Acids

Carboxylic acids have a carbonyl group bonded to a hydroxyl group, and they have the general formula  $R-\overset{\displaystyle O}{\overset{\|}{C}}-O-H$. The functional group,  $-\overset{\displaystyle O}{\overset{\|}{C}}-O-H$, is called the **carboxyl group** (**carbo**nyl + hydro**xyl**):

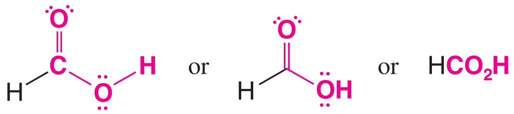


A carboxylic acid

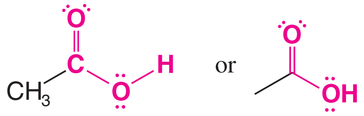


The carboxyl group

Examples of carboxylic acids are formic acid, acetic acid, and benzoic acid:

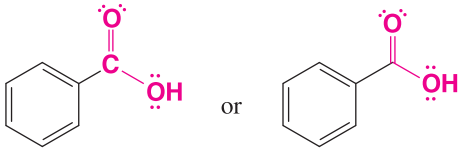


Formic acid



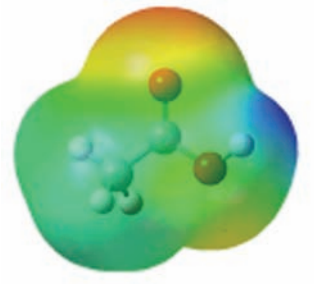
or CH_3CO_2H

Acetic acid



or $C_6H_5CO_2H$

Benzoic acid



Acetic acid

Formic acid is an irritating liquid produced by ants. (The sting of the ant is caused, in part, by formic acid being injected under the skin. *Formic* is the Latin word for ant.) Acetic acid, the substance responsible for the sour taste of vinegar, is produced when certain bacteria act on the ethyl alcohol of wine and cause the ethyl alcohol to be oxidized by air.

Review Problem 2.20

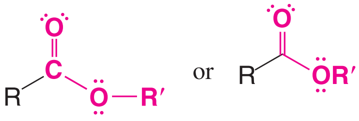
When formic acid donates the proton from its oxygen to a base, a formate ion is the result. Write another resonance structure for formic acid and for the formate ion. Which species, formic acid or the formate ion, would be most stabilized by resonance?

Review Problem 2.21

Write bond-line formulas for four carboxylic acids with the formula $C_5H_{10}O_2$.

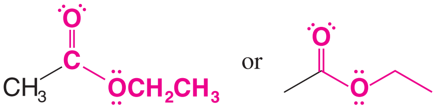
2.10B Esters

Esters have the general formula RCO_2R' (or RCOOR′), where a carbonyl group is bonded to an alkoxyl (—OR) group:



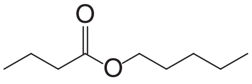
or RCO_2R'

General formula for an ester

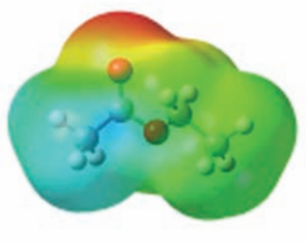


or $CH_3CO_2CH_2CH_3$

Ethyl acetate is an important solvent.



Pentyl butanoate has the odor of apricots and pears.



Ethyl acetate

The ester pentyl butanoate has the odor of apricots and pears.

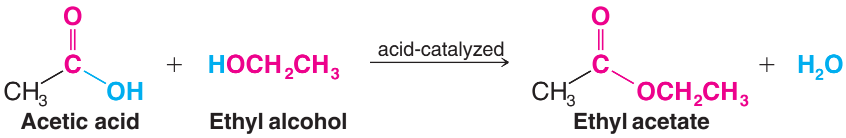
Review Problem 2.22

Write bond-line formulas for three esters with the formula $C_5H_{10}O_2$.

Review Problem 2.23

Write another resonance structure for ethyl acetate. Include formal charges.

Esters can be made from a carboxylic acid and an alcohol through the acid-catalyzed loss of a molecule of water. For example:

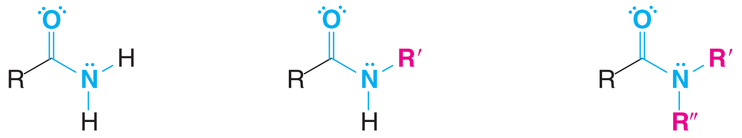


Your body makes esters from long-chain carboxylic acids called “fatty acids” by combining them with glycerol. We discuss their chemistry in detail in Chapter 23.

Nylon is a polymer comprised of regularly repeating amide groups.

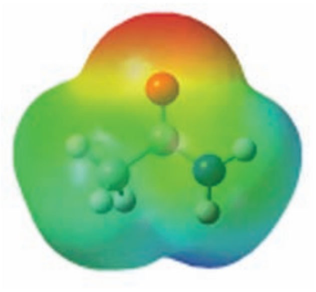
2.10C Amides

Amides have the formulas $RCONH_2$, $RCONHR'$, or $RCONR'R''$ where a carbonyl group is bonded to a nitrogen atom bearing hydrogen and/or alkyl groups. General formulas and some specific examples are shown below.

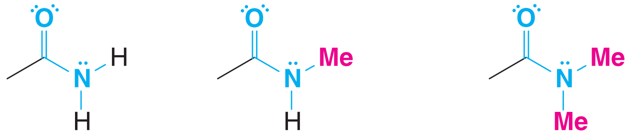


An unsubstituted amide **An *N*-substituted amide** **An *N,N*-disubstituted amide**

General formulas for amides



Acetamide



Acetamide ***N*-Methylacetamide** ***N,N*-Dimethylacetamide**

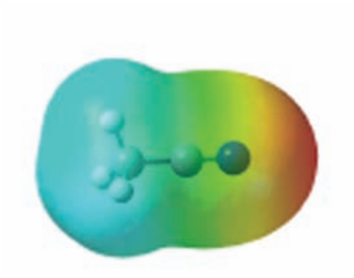
Specific examples of amides

N- and *N,N*- indicate that the substituents are attached to the nitrogen atom.

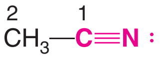
Review Problem 2.24 Write another resonance structure for acetamide.

2.11 Nitriles

A nitrile has the formula R—C≡N: (or R—CN). The carbon and the nitrogen of a nitrile are *sp* hybridized. In IUPAC systematic nomenclature, acyclic nitriles are named by adding the suffix *-nitrile* to the name of the corresponding hydrocarbon. The carbon atom of the —C≡N group is assigned number 1. The name acetonitrile is an acceptable common name for CH_3CN, and acrylonitrile is an acceptable common name for $CH_2{=}CHCN$:

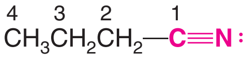


Acetonitrile



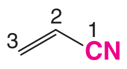
$$\overset{2}{CH_3}{-}\overset{1}{C}{\equiv}N\!:$$

Ethanenitrile (acetonitrile)

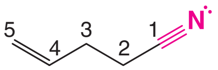


$$\overset{4}{CH_3}\overset{3}{CH_2}\overset{2}{CH_2}{-}\overset{1}{C}{\equiv}N\!:$$

Butanenitrile

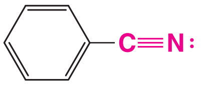


Propenenitrile (acrylonitrile)

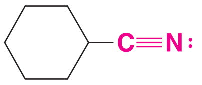


4-Pentenenitrile

Cyclic nitriles are named by adding the suffix *-carbonitrile* to the name of the ring system to which the —CN group is attached. Benzonitrile is an acceptable common name for C_6H_5CN:



Benzenecarbonitrile (benzonitrile)



Cyclohexanecarbonitrile

2.12 Summary of Important Families of Organic Compounds

A summary of the important families of organic compounds is given in Table 2.3. You should learn to identify these common functional groups as they appear in other, more complicated molecules.

2.13 *Physical Properties and Molecular Structure*

So far, we have said little about one of the most obvious characteristics of organic compounds, that is, *their physical state or phase*. Whether a particular substance is a solid, or a liquid, or a gas would certainly be one of the first observations that we would note in any experimental work. The temperatures at which transitions occur between phases, that is, melting points (mp) and boiling points (bp), are also among the more easily measured **physical properties**. Melting points and boiling points are also useful in identifying and isolating organic compounds.

Suppose, for example, we have just carried out the synthesis of an organic compound that is known to be a liquid at room temperature and 1 atm pressure. If we know the boiling point of our desired product and the boiling points of by-products and solvents that may be present in the reaction mixture, we can decide whether or not simple distillation will be a feasible method for isolating our product.

In another instance our product might be a solid. In this case, in order to isolate the substance by crystallization, we need to know its melting point and its solubility in different solvents.

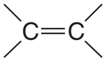
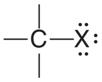
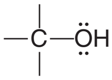
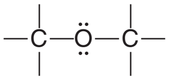
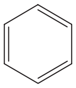
The physical constants of known organic substances are easily found in handbooks and other reference books.* Table 2.4 lists the melting and boiling points of some of the compounds that we have discussed in this chapter.

Often in the course of research, however, the product of a synthesis is a new compound—one that has never been described before. In these instances, success in isolating the new compound depends on making reasonably accurate estimates of its melting point, boiling point, and solubilities. Estimations of these macroscopic physical properties are based on the most likely structure of the substance and on the forces that act between molecules and ions. The temperatures at which phase changes occur are an indication of the strength of these intermolecular forces.

Helpful Hint

Understanding how molecular structure influences physical properties is very useful in practical organic chemistry.

TABLE 2.3 Important Families of Organic Compounds

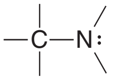
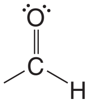
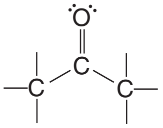
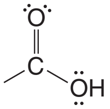
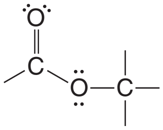
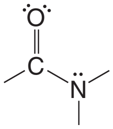
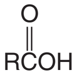
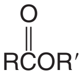
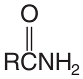
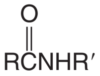
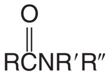
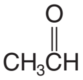
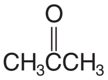
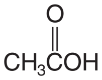
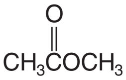
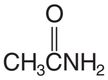
	Family						
	Alkane	**Alkene**	**Alkyne**	**Aromatic**	**Haloalkane**	**Alcohol**	**Ether**
Functional group	C—H and C—C bonds		$-C\equiv C-$	Aromatic ring	 $-\overset{\vert}{\underset{\vert}{C}}-\overset{\cdot\cdot}{\underset{\cdot\cdot}{X}}\colon$	 $-\overset{\vert}{\underset{\vert}{C}}-\overset{\cdot\cdot}{\underset{\cdot\cdot}{O}}H$	 $-\overset{\vert}{\underset{\vert}{C}}-\overset{\cdot\cdot}{\underset{\cdot\cdot}{O}}-\overset{\vert}{\underset{\vert}{C}}-$
General formula	RH	$RCH=CH_2$ $RCH=CHR$ $R_2C=CHR$ $R_2C=CR_2$	$RC\equiv CH$ $RC\equiv CR$	ArH	RX	ROH	ROR
Specific example	CH_3CH_3	$CH_2=CH_2$	$HC\equiv CH$		CH_3CH_2Cl	CH_3CH_2OH	CH_3OCH_3
IUPAC name	Ethane	Ethene	Ethyne	Benzene	Chloroethane	Ethanol	Methoxymethane
Common name[a]	Ethane	Ethylene	Acetylene	Benzene	Ethyl chloride	Ethyl alcohol	Dimethyl ether

[a] These names are also accepted by the IUPAC.

---

*Two useful handbooks are *Handbook of Chemistry*, Lange, N. A., Ed., McGraw-Hill: New York; and *CRC Handbook of Chemistry and Physics*, CRC: Boca Raton, FL.

TABLE 2.3 Important Families of Organic Compounds (*cont.*)

	Family						
	Amine	**Aldehyde**	**Ketone**	**Carboxylic Acid**	**Ester**	**Amide**	**Nitrile**
Functional group							$—C{\equiv}N\!:$
General formula	RNH_2 R_2NH R_3N	 $R\overset{O}{\overset{\|}{C}}H$	 $R\overset{O}{\overset{\|}{C}}R'$	 $R\overset{O}{\overset{\|}{C}}OH$	 $R\overset{O}{\overset{\|}{C}}OR'$	 $R\overset{O}{\overset{\|}{C}}NH_2$  $R\overset{O}{\overset{\|}{C}}NHR'$  $R\overset{O}{\overset{\|}{C}}NR'R''$	RCN
Specific example	CH_3NH_2	 $CH_3\overset{O}{\overset{\|}{C}}H$	 $CH_3\overset{O}{\overset{\|}{C}}CH_3$	 $CH_3\overset{O}{\overset{\|}{C}}OH$	 $CH_3\overset{O}{\overset{\|}{C}}OCH_3$	 $CH_3\overset{O}{\overset{\|}{C}}NH_2$	$CH_3C{\equiv}N$
IUPAC name	Methanamine	Ethanal	Propanone	Ethanoic acid	Methyl ethanoate	Ethanamide	Ethanenitrile
Common name	Methylamine	Acetaldehyde	Acetone	Acetic acid	Methyl acetate	Acetamide	Acetonitrile

2.13A Ionic Compounds: Ion–Ion Forces

The **melting point** of a substance is the temperature at which an equilibrium exists between the well-ordered crystalline state and the more random liquid state. If the substance is an ionic compound, such as sodium acetate (Table 2.4), the **ion–ion forces** that hold the ions together in the crystalline state are the strong electrostatic lattice forces that act between the positive and negative ions in the orderly crystalline structure. In Fig. 2.6 each sodium ion is surrounded by negatively charged acetate ions, and each acetate ion is surrounded by positive sodium ions. A large amount of thermal energy is required to break up the orderly structure of the crystal into the disorderly open structure of a liquid. As a result, the temperature at which sodium acetate melts is quite high, 324°C. The **boiling points** of ionic compounds are higher still, so high that most ionic organic compounds decompose (are changed by undesirable chemical reactions) before they boil. Sodium acetate shows this behavior.

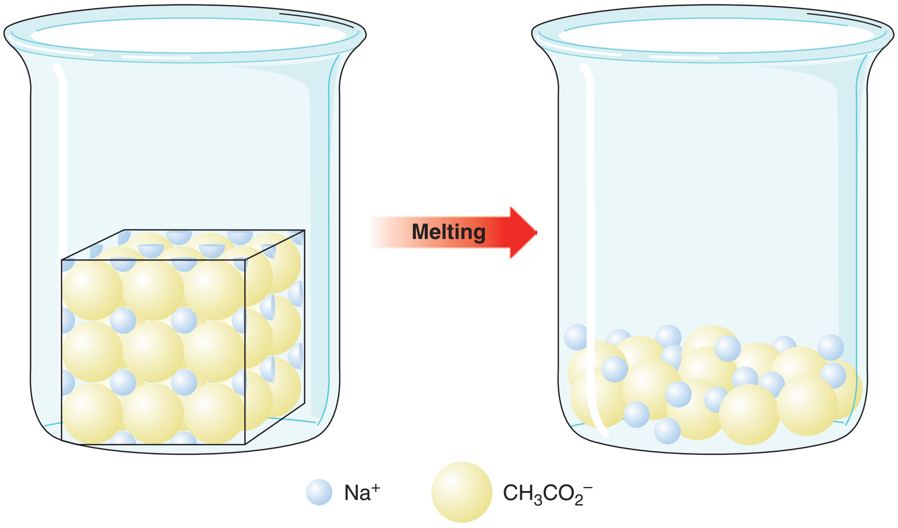
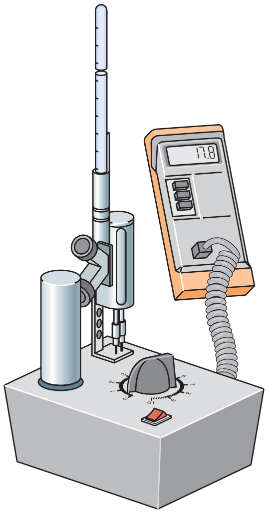


Figure 2.6 The melting of sodium acetate.

TABLE 2.4 Physical Properties of Representative Compounds

Compound	Structure	mp (°C)	bp (°C) (1 atm)[a]
Methane	CH_4	−182.6	−162
Ethane	CH_3CH_3	−172	−88.2
Ethene	$CH_2{=}CH_2$	−169	−102
Ethyne	$HC{\equiv}CH$	−82	−84 subl
Chloromethane	CH_3Cl	−97	−23.7
Chloroethane	CH_3CH_2Cl	−138.7	13.1
Ethyl alcohol	CH_3CH_2OH	−114	78.5
Acetaldehyde	CH_3CHO	−121	20
Acetic acid	CH_3CO_2H	16.6	118
Sodium acetate	CH_3CO_2Na	324	dec
Ethylamine	$CH_3CH_2NH_2$	−80	17
Diethyl ether	$(CH_3CH_2)_2O$	−116	34.6
Ethyl acetate	$CH_3CO_2CH_2CH_3$	−84	77

[a] In this table dec = decomposes and subl = sublimes.



An instrument used to measure melting point

From PAVIA/LAMPMAN/KRIEL. *Introduction to Organic Laboratory Techniques: A Microscale Approach (with Periodic Table)*, 3E. © 1999 Brooks/Cole, a part of Cengage Learning, Inc. Reproduced by permission. www.cengage.com/permissions.

2.13B Intermolecular Forces (van der Waals Forces)

The forces that act between molecules are not as strong as those between ions, but they account for the fact that even completely nonpolar molecules can exist in liquid and solid states. These intermolecular forces, collectively called **van der Waals forces**, are all electrical in nature. We will focus our attention on three types:

1. Dipole–dipole forces
2. Hydrogen bonds
3. Dispersion forces

Dipole–Dipole Forces Most organic molecules are not fully ionic but have instead a *permanent dipole moment* resulting from a nonuniform distribution of the bonding electrons (Section 2.3). Acetone and acetaldehyde are examples of molecules with permanent dipoles because the carbonyl group that they contain is highly polarized. In these compounds, the attractive forces between molecules are much easier to visualize. In the liquid or solid state, **dipole–dipole** attractions cause the molecules to orient themselves so that the positive end of one molecule is directed toward the negative end of another (Fig. 2.7).

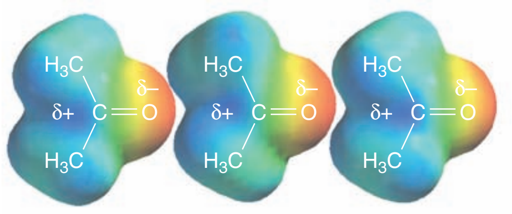


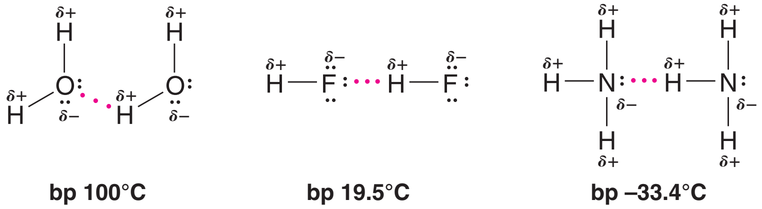
Figure 2.7 Electrostatic potential models for acetone molecules that show how acetone molecules might align according to attractions of their partially positive regions and partially negative regions (dipole–dipole interactions).

Hydrogen Bonds

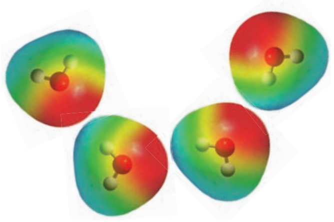
- Very strong dipole–dipole attractions occur between hydrogen atoms bonded to small, strongly electronegative atoms (O, N, or F) and nonbonding electron pairs on other such electronegative atoms. This type of intermolecular force is called a **hydrogen bond**.

Hydrogen bonds (bond dissociation energies of about 4–38 kJ mol^{-1}) are weaker than ordinary covalent bonds but much stronger than the dipole–dipole interactions that occur above, for example, in acetone.

Hydrogen bonding explains why water, ammonia, and hydrogen fluoride all have far higher boiling points than methane (bp −161.6°C), even though all four compounds have similar molecular weights.



Hydrogen bonds are shown by the red dots.

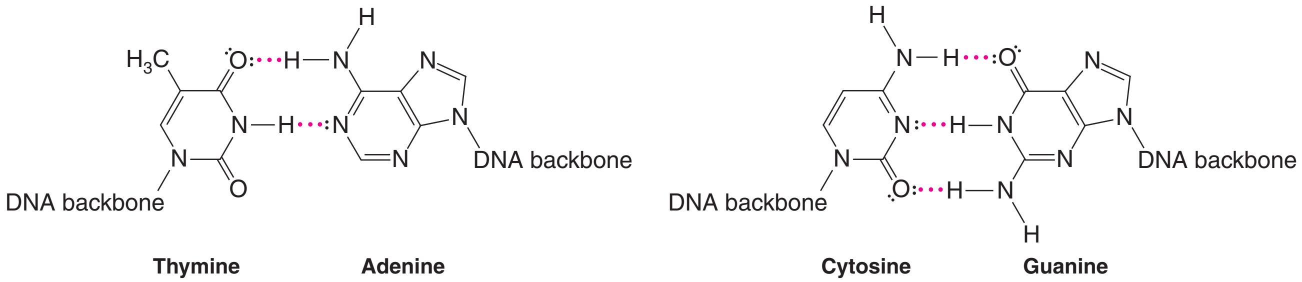


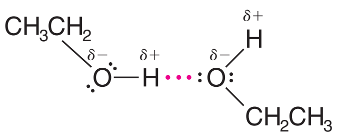
Water molecules associated by attraction of opposite partial charges.

One of the most important consequences of hydrogen bonding is that it causes water to be a liquid rather than a gas at 25°C. Calculations indicate that in the absence of hydrogen bonding, water would have a boiling point near −80°C and would not exist as a liquid unless the temperature were lower than that temperature. Had this been the case, it is highly unlikely that life, as we know it, could have developed on the planet Earth.

Hydrogen bonds hold the base pairs of double-stranded DNA together (see Section 25.4). Thymine hydrogen bonds with adenine. Cytosine hydrogen bonds with guanine.

Hydrogen bonding accounts for the fact that ethyl alcohol has a much higher boiling point (+78.5°C) than dimethyl ether (−24.9°C) even though the two compounds have the same molecular weight. Molecules of ethyl alcohol, because they have a hydrogen atom covalently bonded to an oxygen atom, can form strong hydrogen bonds to each other.



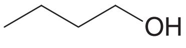
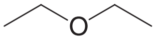
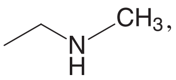


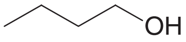
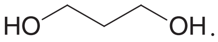
The red dots represent a hydrogen bond. Strong hydrogen bonding is limited to molecules having a hydrogen atom attached to an O, N, or F atom.

Molecules of dimethyl ether, because they lack a hydrogen atom attached to a strongly electronegative atom, cannot form strong hydrogen bonds to each other. In dimethyl ether the intermolecular forces are weaker dipole–dipole interactions.

Review Problem 2.25

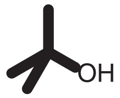
The compounds in each part below have the same (or similar) molecular weights. Which compound in each part would you expect to have the higher boiling point? Explain your answers.

(a)  or , **(b)** $(CH_3)_3N$ or ,

(c)  or .

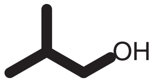
A factor (in addition to polarity and hydrogen bonding) that affects the *melting point* of many organic compounds is the compactness and rigidity of their individual molecules.

- Molecules that are symmetrical generally have abnormally high melting points. *tert*-Butyl alcohol, for example, has a much higher melting point than the other isomeric alcohols shown here:

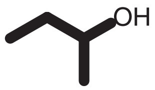


***tert*-Butyl alcohol**
(mp 25°C)

Butyl alcohol
(mp −90°C)



Isobutyl alcohol
(mp −108°C)



***sec*-Butyl alcohol**
(mp −114°C)

Review Problem 2.26

Which compound would you expect to have the higher melting point, propane or cyclopropane? Explain your answer.

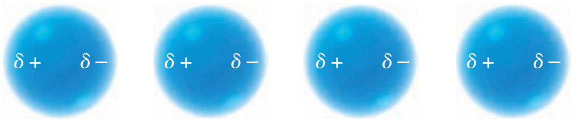
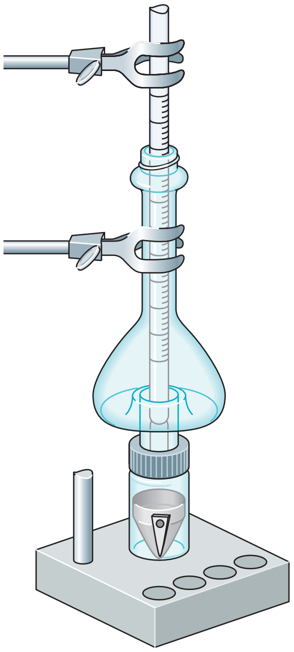


Figure 2.8 Temporary dipoles and induced dipoles in nonpolar molecules resulting from an uneven distribution of electrons at a given instant.

Dispersion Forces If we consider a substance like methane where the particles are nonpolar molecules, we find that the melting point and boiling point are very low: −182.6°C and −162°C, respectively. Instead of asking, "Why does methane melt and boil at low temperatures?" a more appropriate question might be "Why does methane, a nonionic, nonpolar substance, become a liquid or a solid at all?" The answer to this question can be given in terms of attractive intermolecular forces called **dispersion forces** or London forces.

An accurate account of the nature of dispersion forces requires the use of quantum mechanics. We can, however, visualize the origin of these forces in the following way. The average distribution of charge in a nonpolar molecule (such as methane) over a period of time is uniform. At any given instant, however, *because electrons move*, the electrons and therefore the charge may not be uniformly distributed. Electrons may, in one instant, be slightly accumulated on one part of the molecule, and, as a consequence, *a small temporary dipole will occur* (Fig. 2.8). This temporary dipole in one molecule can induce opposite (attractive) dipoles in surrounding molecules. It does this because the negative (or positive) charge in a portion of one molecule will distort the electron cloud of an adjacent portion of another molecule, causing an opposite charge to develop there. These temporary dipoles change constantly, but the net result of their existence is to produce attractive forces between nonpolar molecules and thus make possible the existence of their liquid and solid states.

Two important factors determine the magnitude of dispersion forces.



A microscale distillation apparatus

From PAVIA/LAMPMAN/KRIEL. *Introduction to Organic Laboratory Techniques: A Microscale Approach (with Periodic Table)*, 3E. © 1999 Brooks/Cole, a part of Cengage Learning, Inc. Reproduced by permission. www.cengage.com/permissions.

1. **The relative polarizability of electrons of the atoms involved**. *By polarizability we mean how easily the electrons respond to a changing electric field*. The electrons of large atoms such as iodine are loosely held and are easily polarized, while the electrons of small atoms such as fluorine are more tightly held and are much less polarizable. Atoms with unshared pairs are more easily polarized than atoms with only bonding pairs. Table 2.5 gives the relative magnitudes of dispersion forces and dipole–dipole interactions for several simple compounds. Notice with HI the dispersion forces are far more important than dipole–dipole forces, whereas with H_2O, dipole–dipole forces (of the kind we call hydrogen bonds) are more important.

TABLE 2.5 Attractive Energies in Simple Covalent Compounds

		Attractive Energies (kJ mol^{-1})			
Molecule	Dipole Moment (D)	Dipole–Dipole	Dispersion	Melting Point (°C)	Boiling Point (°C) at 1 atm
H_2O	1.85	36[a]	8.8	0	100
NH_3	1.47	14[a]	15	−78	−33
HCl	1.08	3[a]	17	−115	−85
HBr	0.80	0.8	22	−88	−67
HI	0.42	0.03	28	−51	−35

[a] These dipole–dipole attractions are called hydrogen bonds.

Dispersion forces are what provides a gecko's grip to smooth surfaces.

CF_4 and CI_4 are both nonpolar molecules. But if we were to consider the intermolecular forces between two CI_4 molecules, which contain polarizable iodine atoms, we would find that the dispersion forces are much larger than between two CF_4 molecules, which contains fluorine atoms that are not very polarizable.

2. **The relative surface area of the molecules involved**. The larger the surface area, the larger is the overall attraction between molecules caused by dispersion forces. Molecules that are generally longer, flatter, or cylindrical have a greater surface area available for intermolecular interactions than more spherical molecules, and consequently have greater attractive forces between them than the tangential interactions between branched molecules. This is evident when comparing pentane, the unbranched C_5H_{12} hydrocarbon, with neopentane, the most highly branched C_5H_{12} isomer (in which one carbon bears four methyl groups). Pentane has a boiling point of 36.1°C. Neopentane has a boiling point of 9.5°C. The difference in their boiling points indicate that the attractive forces between pentane molecules are stronger than between neopentane molecules.

For large molecules, the cumulative effect of these small and rapidly changing dispersion forces can lead to a large net attraction.

2.13C Boiling Points

The *boiling point* of a liquid is the temperature at which the vapor pressure of the liquid equals the pressure of the atmosphere above it. For this reason, the boiling points of liquids are *pressure dependent*, and boiling points are always reported as occurring at a particular pressure, at 1 atm (or at 760 torr), for example. A substance that boils at 150°C at 1 atm pressure will boil at a substantially lower temperature if the pressure is reduced to, for example, 0.01 torr (a pressure easily obtained with a vacuum pump). The normal boiling point given for a liquid is its boiling point at 1 atm.

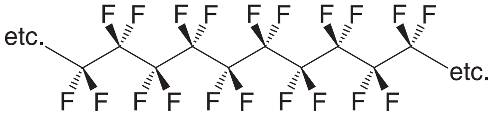
In passing from a liquid to a gaseous state, the individual molecules (or ions) of the substance must separate. Because of this, we can understand why ionic organic compounds often decompose before they boil. The thermal energy required to completely separate (volatilize) the ions is so great that chemical reactions (decompositions) occur first.

THE CHEMISTRY OF . . .

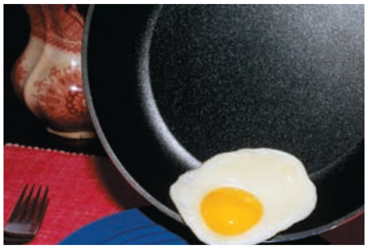
Fluorocarbons and Teflon

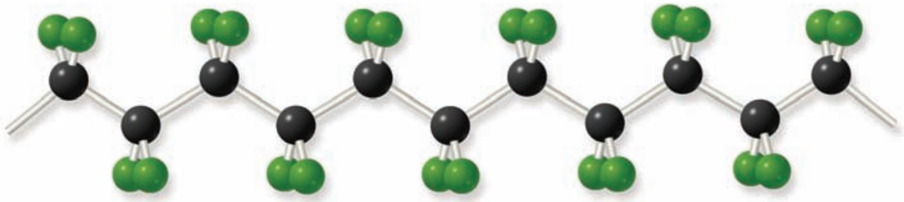
Fluorocarbons (compounds containing only carbon and fluorine) have extraordinarily low boiling points when compared to hydrocarbons of the same molecular weight. The fluorocarbon C_5F_{12} has a slightly lower boiling point than pentane (C_5H_{12}) even though it has a far higher molecular weight. The important factor in explaining this behavior is the very low polarizability of fluorine atoms that we mentioned earlier, resulting in very small dispersion forces.

The fluorocarbon called *Teflon* $[CF_2CF_2]_n$ (see Section 10.10) has self-lubricating properties that are exploited in making "nonstick" frying pans and lightweight bearings.



Teflon

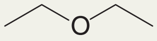




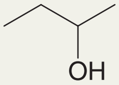
Nonpolar compounds, where the intermolecular forces are very weak, usually boil at low temperatures even at 1 atm pressure. This is not always true, however, because of other factors that we have not yet mentioned: the effects of molecular weight and molecular shape and surface area. Heavier molecules require greater thermal energy in order to acquire velocities sufficiently great to escape the liquid phase, and because the surface areas of larger molecules can be much greater, intermolecular dispersion attractions can also be much larger. These factors explain why nonpolar ethane (bp −88.2°C) boils higher than methane (bp −162°C) at a pressure of 1 atm. It also explains why, at 1 atm, the even heavier and larger nonpolar molecule decane ($C_{10}H_{22}$) boils at +174°C. The relationship between dispersion forces and surface area helps us understand why neopentane (2,2-dimethylpropane) has a lower boiling point (9.5°C) than pentane (36.1°C), even though they have the same molecular weight. The branched structure of neopentane allows less surface interaction between neopentane molecules, hence lower dispersion forces, than does the linear structure of pentane.

Solved Problem 2.6

Arrange the following compounds according to their expected boiling points, with the lowest boiling point first, and explain your answer. Notice that the compounds have similar molecular weights.



Diethyl ether



***sec*-Butyl alcohol**

Pentane

STRATEGY AND ANSWER

Pentane < diethyl ether < *sec*-butyl alcohol

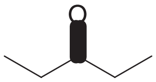
Increasing boiling point →

Pentane has no polar groups and has only dispersion forces holding its molecules together. It would have the lowest boiling point. Diethyl ether has the polar ether group that provides dipole–dipole forces which are greater than dispersion forces, meaning it would have a higher boiling point than pentane. *sec*-Butyl alcohol has an —OH group that can form strong hydrogen bonds; therefore, it would have the highest boiling point.

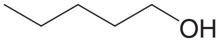
Review Problem 2.27

Arrange the following compounds in order of increasing boiling point. Explain your answer in terms of the intermolecular forces in each compound.

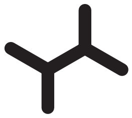
(a)



(b)



(c)



(d)

2.13D Solubilities

Intermolecular forces are of primary importance in explaining the **solubilities** of substances. Dissolution of a solid in a liquid is, in many respects, like the melting of a solid. The orderly crystal structure of the solid is destroyed, and the result is the formation of the more disorderly arrangement of the molecules (or ions) in solution. In the process of dissolving, too, the molecules or ions must be separated from each other, and energy must be supplied for both changes. The energy required to overcome lattice energies and intermolecular or interionic attractions comes from the formation of new attractive forces between solute and solvent.

Helpful Hint

Your ability to make qualitative predictions regarding solubility will prove very useful in the organic chemistry laboratory.

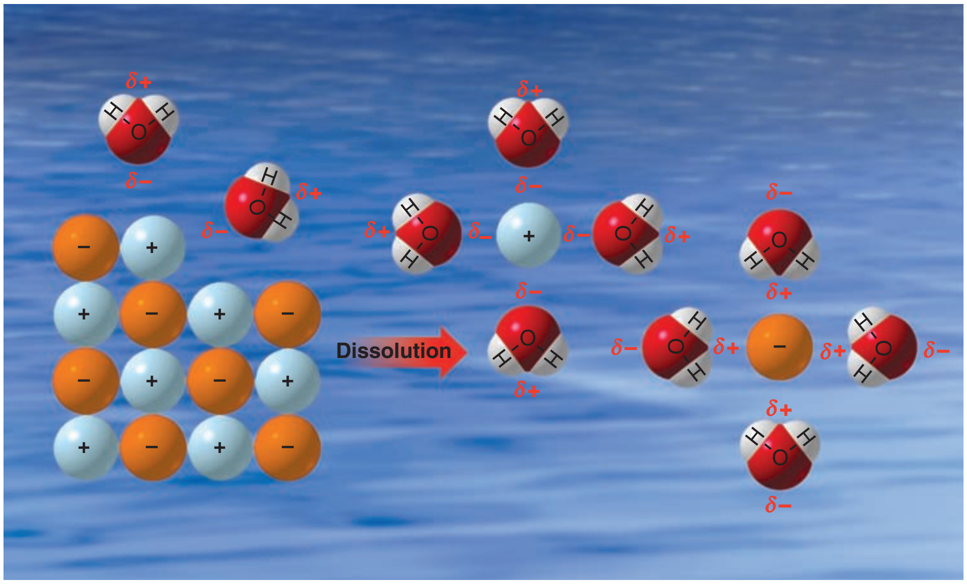


Figure 2.9 The dissolution of an ionic solid in water, showing the hydration of positive and negative ions by the very polar water molecules. The ions become surrounded by water molecules in all three dimensions, not just the two shown here.

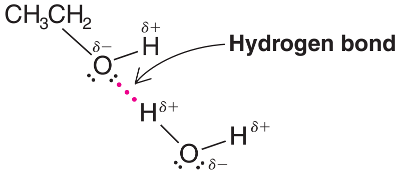
Consider the dissolution of an ionic substance as an example. Here both the lattice energy and interionic attractions are large. We find that water and only a few other very polar solvents are capable of dissolving ionic compounds. These solvents dissolve ionic compounds by **hydrating** or **solvating** the ions (Fig. 2.9).

Water molecules, by virtue of their great polarity as well as their very small, compact shape, can very effectively surround the individual ions as they are freed from the crystal surface. Positive ions are surrounded by water molecules with the negative end of the water dipole pointed toward the positive ion; negative ions are solvated in exactly the opposite way. Because water is highly polar, and because water is capable of forming strong hydrogen bonds, the **ion–dipole forces** of attraction are also large. The energy supplied by the formation of these forces is great enough to overcome both the lattice energy and interionic attractions of the crystal.

A general rule for solubility is that "like dissolves like" in terms of comparable polarities.

- Polar and ionic solids are usually soluble in polar solvents.
- Polar liquids are usually miscible.
- Nonpolar solids are usually soluble in nonpolar solvents.
- Nonpolar liquids are usually miscible.
- Polar and nonpolar liquids, like oil and water, are usually not soluble to large extents.

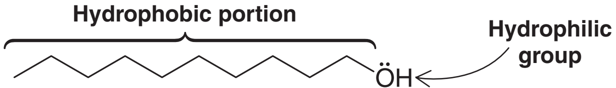
Methanol and water are miscible in all proportions; so too are mixtures of ethanol and water and mixtures of both propyl alcohols and water. In these cases the alkyl groups of the alcohols are relatively small, and the molecules therefore resemble water more than they do an alkane. Another factor in understanding their solubility is that the molecules are capable of forming strong hydrogen bonds to each other:



We often describe molecules or parts of molecules as being hydrophilic or hydrophobic. The alkyl groups of methanol, ethanol, and propanol are hydrophobic. Their hydroxyl groups are hydrophilic.

- **Hydrophobic** means incompatible with water (*hydro*, water; *phobic*, fearing or avoiding).
- **Hydrophilic** means compatible with water (*philic*, loving or seeking).

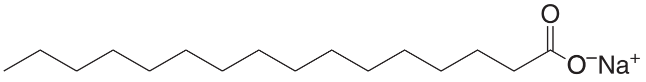
Decyl alcohol, with a chain of 10 carbon atoms, is a compound whose hydrophobic alkyl group overshadows its hydrophilic hydroxyl group in terms of water solubility.



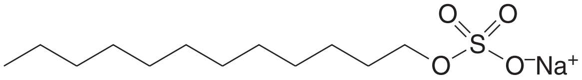
Decyl alcohol

An explanation for why nonpolar groups such as long alkane chains avoid an aqueous environment, that is, for the so-called **hydrophobic effect**, is complex. The most important factor seems to involve an **unfavorable entropy change** in the water. Entropy changes (Section 3.10) have to do with changes from a relatively ordered state to a more disordered one or the reverse. Changes from order to disorder are favorable, whereas changes from disorder to order are unfavorable. For a nonpolar hydrocarbon chain to be accommodated by water, the water molecules have to form a more ordered structure around the chain, and for this, the entropy change is unfavorable.

We will see in Section 23.2C that the presence of a hydrophobic group and a hydrophilic group are essential components of soaps and detergents.



A typical soap molecule



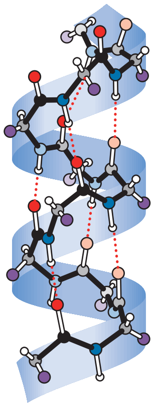
A typical detergent molecule

The hydrophobic long carbon chains of a soap or detergent embed themselves in the oily layer that typically surrounds the thing we want to wash away. The hydrophilic ionic groups at the ends of the chains are then left exposed on the surface and make the surface one that water molecules find attractive. Oil and water don't mix, but now the oily layer looks like something ionic and the water can take it "right down the drain."

2.13E Guidelines for Water Solubility

Organic chemists usually define a compound as water soluble if at least 3 g of the organic compound dissolves in 100 mL of water. We find that for compounds containing one hydrophilic group—and thus capable of forming strong hydrogen bonds—the following approximate guidelines hold: Compounds with one to three carbon atoms are water soluble, compounds with four or five carbon atoms are borderline, and compounds with six carbon atoms or more are insoluble.

When a compound contains more than one hydrophilic group, these guidelines do not apply. Polysaccharides (Chapter 22), proteins (Chapter 24), and nucleic acids (Chapter 25) all contain thousands of carbon atoms *and many are water soluble*. They dissolve in water because they also contain thousands of hydrophilic groups.



Hydrogen bonding (red dotted lines) in the α-helix structure of proteins

(Illustration, Irving Geis. Rights owned by Howard Hughes Medical Institute. Not to be reproduced without permission.)

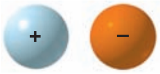
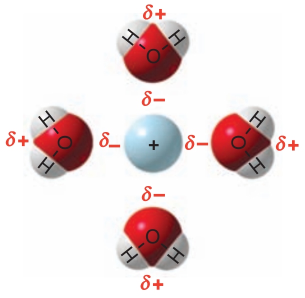
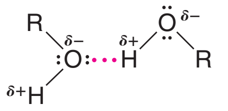
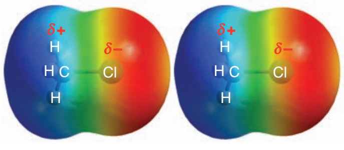
2.13F Intermolecular Forces in Biochemistry

Later, after we have had a chance to examine in detail the properties of the molecules that make up living organisms, we shall see how **intermolecular forces** are extremely important in the functioning of cells. **Hydrogen bond** formation, the hydration of polar groups, and the tendency of nonpolar groups to avoid a polar environment all cause complex protein molecules to fold in precise ways—ways that allow them to function as biological catalysts of incredible efficiency. The same factors allow molecules of hemoglobin to assume the shape needed to transport oxygen. They allow proteins and molecules called lipids to function as cell membranes. Hydrogen bonding gives certain carbohydrates a globular shape that makes them highly efficient food reserves in animals. It gives molecules of other carbohydrates a rigid linear shape that makes them perfectly suited to be structural components in plants.

2.14 *Summary of Attractive Electric Forces*

The attractive forces occurring between molecules and ions that we have studied so far are summarized in Table 2.6.

TABLE 2.6 Attractive Electric Forces

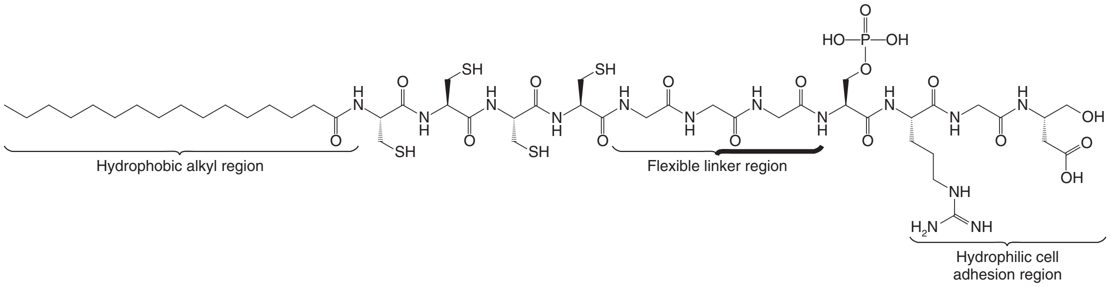
Electric Force	Relative Strength	Type	Example
Cation–anion (in a crystal)	Very strong		Sodium chloride crystal lattice
Covalent bonds	Strong (140–523 kJ mol^{-1})	Shared electron pairs	H—H (436 kJ mol^{-1}) CH_3—CH_3 (378 kJ mol^{-1}) I—I (151 kJ mol^{-1})
Ion–dipole	Moderate		Na^+ in water (see Fig. 2.9)
Hydrogen bonds	Moderate to weak (4–38 kJ mol^{-1})	$-\overset{\delta-}{Z}:\cdots\overset{\delta+}{H}-$	
Dipole–dipole	Weak	$\overset{\delta+}{CH_3}\overset{\delta-}{Cl}\cdots\overset{\delta+}{CH_3}\overset{\delta-}{Cl}$	
Dispersion	Variable	Transient dipole	Interactions between methane molecules

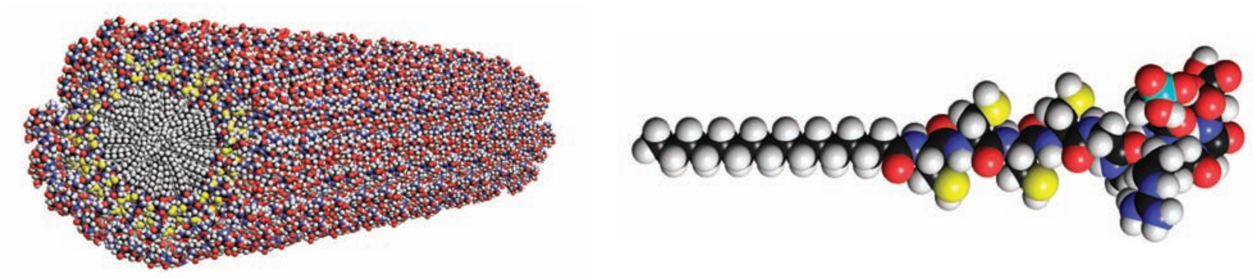
THE CHEMISTRY OF . . .

Organic Templates Engineered to Mimic Bone Growth

Intermolecular forces play a myriad of roles in life and in the world around us. Intermolecular forces hold together the strands of our DNA, provide structure to our cell membranes, cause the feet of gecko lizards to stick to walls and ceilings, keep water from boiling at room temperature and ordinary pressure, and literally provide the adhesive forces that hold our cells, bones, and tissues together. As these examples show, the world around us provides exquisite instruction in nanotechnology and bioengineering, and scientists throughout the ages have been inspired to create and innovate based on nature. One target of recent research in bioengineering is the development of synthetic materials that mimic nature's template for bone growth. A synthetic material with bone-promoting properties could be used to help repair broken bones, offset osteoporosis, and treat bone cancer.

Both natural bone growth and the synthetic system under development depend strongly on intermolecular forces. In living systems, bones grow by adhesion of specialized cells to a long fibrous natural template called collagen. Certain functional groups along the collagen promote the binding of bone-growing cells, while other functional groups facilitate calcium crystallization. Chemists at Northwestern University (led by S. I. Stupp) have engineered a molecule that can be made in the laboratory and that mimics this process. The molecule shown below spontaneously self-assembles into a long tubular aggregate, imitating the fibers of collagen. Dispersion forces between hydrophobic alkyl tails on the molecule cause self-assembly of the molecules into tubules. At the other end of the molecule, the researchers included functional groups that promote cell binding and still other functional groups that encourage calcium crystallization. Lastly, they included functional groups that allow one molecule to be covalently linked to its neighbors after the self-assembly process has occurred, thus adding further stabilization to the initially noncovalent structure. Designing all of these features into the molecular structure has paid off, because the self-assembled fiber promotes calcium crystallization along its axis, much like nature's collagen template. This example of molecular design is just one exciting development at the intersection of nanotechnology and bioengineering.





(Reprinted with permission from Hartgerink, J. D., Beniash, E., Stupp, S. I., SCIENCE 294: 1684–1688, Figure 1 (2001). Copyright 2001 AAAS.)

2.15 *Infrared Spectroscopy: An Instrumental Method for Detecting Functional Groups*

Infrared (IR) spectroscopy is a simple and rapid instrumental technique that can give evidence for the presence of various functional groups. If you had a sample of unknown identity, among the first things you would do is obtain an infrared spectrum, along with determining its solubility in common solvents and its melting and/or boiling point.

Infrared spectroscopy, as all forms of spectroscopy, depends on the interaction of molecules or atoms with electromagnetic radiation. Infrared radiation causes atoms and groups of atoms of organic compounds to vibrate with increased amplitude about the covalent bonds that connect them. (Infrared radiation is not of sufficient energy to excite electrons, as is the case when some molecules interact with visible, ultraviolet, or higher energy forms of light.) Since the functional groups of organic molecules include specific arrangements of bonded atoms, absorption of IR radiation by an organic molecule will occur at specific frequencies characteristic of the types of bonds and atoms present in the specific functional groups of that molecule. These vibrations are *quantized*, and as they occur, the compounds absorb IR energy in particular regions of the IR portion of the spectrum.

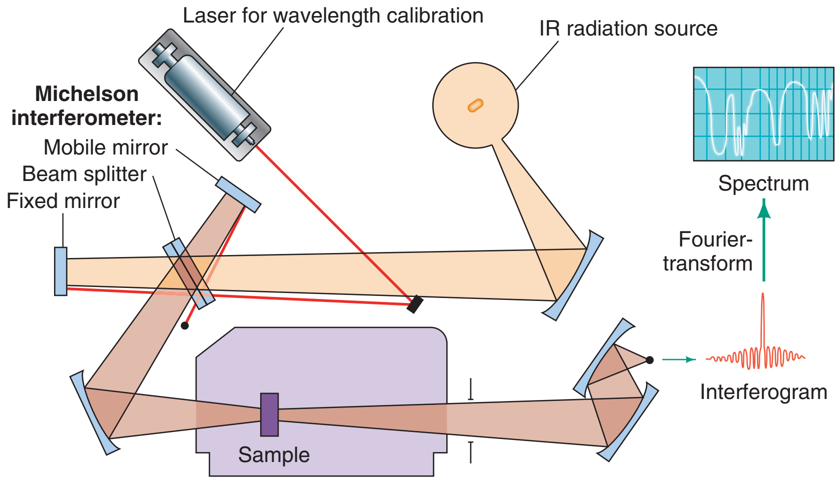


Figure 2.10 A diagram of a Fourier transform infrared (FTIR) spectrometer. FTIR spectrometers employ a Michelson interferometer, which splits the radiation beam from the IR source so that it reflects simultaneously from a moving mirror and a fixed mirror, leading to interference. After the beams recombine, they pass through the sample to the detector and are recorded as a plot of time versus signal intensity, called an interferogram. The overlapping wavelengths and the intensities of their respective absorptions are then converted to a spectrum by applying a mathematical operation called a Fourier transform.

The FTIR method eliminates the need to scan slowly over a range of wavelengths, as was the case with older types of instruments called dispersive IR spectrometers, and therefore FTIR spectra can be acquired very quickly. The FTIR method also allows greater throughput of IR energy. The combination of these factors gives FTIR spectra strong signals as compared to background noise (i.e., a high signal to noise ratio) because radiation throughput is high and rapid scanning allows multiple spectra to be averaged in a short period of time. The result is enhancement of real signals and cancellation of random noise. (Diagram adapted from the computer program IR Tutor, Columbia University.)

An infrared spectrometer (Fig. 2.10) operates by passing a beam of IR radiation through a sample and comparing the radiation transmitted through the sample with that transmitted in the absence of the sample. Any frequencies absorbed by the sample will be apparent by the difference. The spectrometer plots the results as a graph showing absorbance versus frequency or wavelength.

- The position of an absorption band (peak) in an IR spectrum is specified in units of **wavenumbers ($\bar{\nu}$)**.

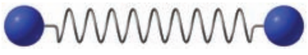
Wavenumbers are the reciprocal of wavelength when wavelength is expressed in centimeters (the unit is cm^{-1}), and therefore give the number of wave cycles per centimeter. The larger the wavenumber, the higher is the frequency of the wave, and correspondingly the higher is the frequency of the bond absorption. IR absorptions are sometimes, though less commonly, reported in terms of **wavelength (λ)**, in which case the units are micrometers (μm; old name micron, μ). Wavelength is the distance from crest to crest of a wave.

$$\bar{\nu} = \frac{1}{\lambda} \quad \text{(with } \lambda \text{ in cm)} \qquad \text{or} \qquad \bar{\nu} = \frac{10{,}000}{\lambda} \quad \text{(with } \lambda \text{ in } \mu\text{m)}$$

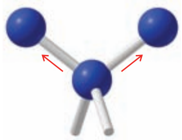
In their vibrations covalent bonds behave as if they were tiny springs connecting the atoms. When the atoms vibrate, they can do so only at certain frequencies, as if the bonds were "tuned." Because of this, covalently bonded atoms have only particular vibrational energy levels; that is, the levels are quantized.

The excitation of a molecule from one vibrational energy level to another occurs only when the compound absorbs IR radiation of a particular energy, meaning a particular wavelength or frequency. Note that the energy (E) of absorption is directly proportional to the **frequency** of radiation (ν) because $\Delta E = h\nu$, and inversely proportional to the wavelength (λ) because $\nu = \dfrac{c}{\lambda}$, and therefore $\Delta E = \dfrac{hc}{\lambda}$.

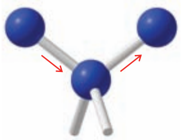
Molecules can vibrate in a variety of ways. Two atoms joined by a covalent bond can undergo a stretching vibration where the atoms move back and forth as if joined by a spring. Three atoms can also undergo a variety of stretching and bending vibrations.



A stretching vibration

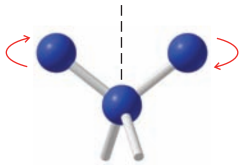


Symmetric stretching



Asymmetric stretching

An in-plane bending vibration (scissoring)



An out-of-plane bending vibration (twisting)

The *frequency* of a given stretching vibration *in an IR spectrum* can be related to two factors. These are *the masses of the bonded atoms*—light atoms vibrate at higher frequencies than heavier ones—*and the relative stiffness of the bond*. (These factors are accounted for in Hooke's law, a relationship you may study in introductory physics.) Triple bonds are stiffer (and vibrate at higher frequencies) than double bonds, and double bonds are stiffer (and vibrate at higher frequencies) than single bonds. We can see some of these effects in Table 2.7. Notice that stretching frequencies of groups involving hydrogen (a light atom) such as C—H, N—H, and O—H all occur at relatively high frequencies:

GROUP	BOND	FREQUENCY RANGE (cm^{-1})
Alkyl	C—H	2853–2962
Alcohol	O—H	3590–3650
Amine	N—H	3300–3500

Notice, too, that triple bonds vibrate at higher frequencies than double bonds:

BOND	FREQUENCY RANGE (cm^{-1})
C≡C	2100–2260
C≡N	2220–2260
C═C	1620–1680
C═O	1630–1780

- Not all molecular vibrations result in the absorption of IR energy. ***In order for a vibration to occur with the absorption of IR energy, the dipole moment of the molecule must change as the vibration occurs***.

Thus, methane does not absorb IR energy for symmetric streching of the four C—H bonds; asymmetric stretching, on the other hand, does lead to an IR absorption. Symmetrical vibrations of the carbon–carbon double and triple bonds of ethene and ethyne do not result in the absorption of IR radiation, either.

Vibrational absorption may occur outside the region measured by a particular IR spectrometer, and vibrational absorptions may occur so closely together that peaks fall on top of peaks.

TABLE 2.7 Characteristic Infrared Absorptions of Groups

Group		Frequency Range ($\mathrm{cm^{-1}}$)	Intensity[a]
A. Alkyl			
C—H (stretching)		2853–2962	(m–s)
Isopropyl, $-\mathrm{CH(CH_3)_2}$		1380–1385	(s)
	and	1365–1370	(s)
tert-Butyl, $-\mathrm{C(CH_3)_3}$		1385–1395	(m)
	and	~1365	(s)
B. Alkenyl			
C—H (stretching)		3010–3095	(m)
$\mathrm{C{=}C}$ (stretching)		1620–1680	(v)
$\mathrm{R{-}CH{=}CH_2}$ (out-of-plane C—H bendings)		985–1000	(s)
	and	905–920	(s)
$\mathrm{R_2C{=}CH_2}$		880–900	(s)
cis-$\mathrm{RCH{=}CHR}$		675–730	(s)
trans-$\mathrm{RCH{=}CHR}$		960–975	(s)
C. Alkynyl			
$\mathrm{{\equiv}C{-}H}$ (stretching)		~3300	(s)
$\mathrm{C{\equiv}C}$ (stretching)		2100–2260	(v)
D. Aromatic			
Ar—H (stretching)		~3030	(v)
C$\cdots$C (stretching)		1450–1600	(m)
Aromatic substitution type (C—H out-of-plane bendings)			
Monosubstituted		690–710	(very s)
	and	730–770	(very s)
o-Disubstituted		735–770	(s)
m-Disubstituted		680–725	(s)
	and	750–810	(very s)
p-Disubstituted		800–860	(very s)
E. Alcohols, Phenols, and Carboxylic Acids			
O—H (stretching)			
Alcohols, phenols (dilute solutions)		3590–3650	(sharp, v)
Alcohols, phenols (hydrogen bonded)		3200–3550	(broad, s)
Carboxylic acids (hydrogen bonded)		2500–3000	(broad, v)
F. Ethers and Alcohols			
C—O—C (stretching)		1020–1275	(s)
G. Aldehydes, Ketones, Ethers, Carboxylic Acids, and Amides			
$\mathrm{C{=}O}$ (stretching)		1630–1780	(s)
Aldehydes		1690–1740	(s)
Ketones		1680–1750	(s)
Esters		1735–1750	(s)
Carboxylic acids		1710–1780	(s)
Amides		1630–1690	(s)
H. Amines			
N—H		3300–3500	(m)
I. Nitriles			
$\mathrm{C{\equiv}N}$		2220–2260	(m)

[a] Abbreviations: s = strong, m = medium, w = weak, v = variable, ~ = approximately.

Other factors bring about even more absorption peaks. Overtones (harmonics) of fundamental absorption bands may be seen in IR spectra even though these overtones occur with greatly reduced intensity. Bands called combination bands and difference bands also appear in IR spectra.

Because IR spectra of even relatively simple compounds contain so many peaks, the possibility that two different compounds will have the same IR spectrum is exceedingly small. It is because of this that an IR spectrum has been called the "fingerprint" of a molecule. Thus, with organic compounds, if two pure samples give different IR spectra, one can be certain that they are different compounds. If they give the same IR spectrum, then they are very likely to be the same compound.

2.16 *Interpreting IR Spectra*

IR spectra contain a wealth of information about the structures of compounds. We show some of the information that can be gathered from the spectra of octane and methylbenzene (commonly called toluene) in Figs. 2.11 and 2.12. In this section we shall learn how to recognize the presence of characteristic IR absorption peaks that result from vibrations of alkyl and functional groups. The data given in Table 2.7 will provide us with key information to use when correlating actual spectra with IR absorption frequencies that are typical for various groups.

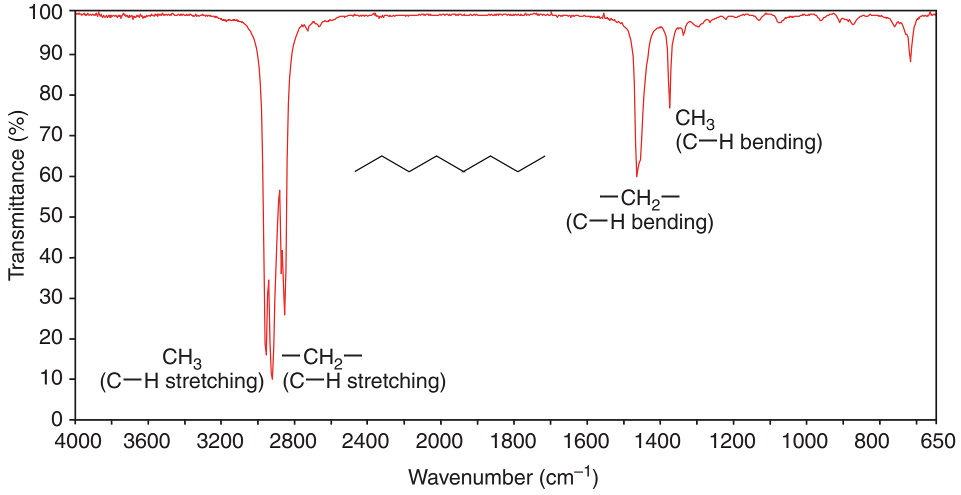


Figure 2.11 The IR spectrum of octane. (Notice that, in IR spectra, the peaks are usually measured in % transmittance. Thus, the peak at 2900 cm^{-1} has 10% transmittance, that is, an absorbance, *A*, of 0.90.)

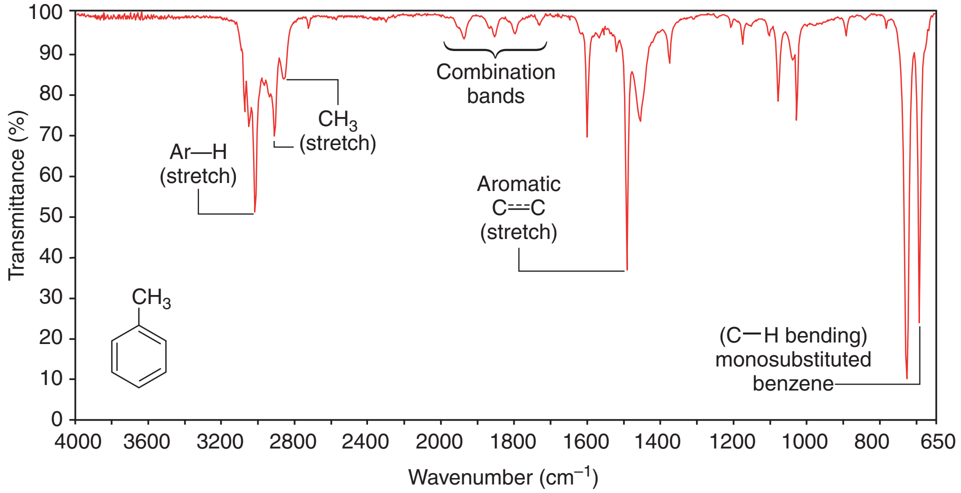


Figure 2.12 The IR spectrum of methylbenzene (toluene).

2.16A Infrared Spectra of Hydrocarbons

- All hydrocarbons give absorption peaks in the 2800–3300-cm^{-1} region that are associated with carbon–hydrogen stretching vibrations.

We can use these peaks in interpreting IR spectra because the exact location of the peak depends on the strength (and stiffness) of the C—H bond, which in turn depends on the hybridization state of the carbon that bears the hydrogen. The C—H bonds involving *sp*-hybridized carbon are strongest and those involving sp^3-hybridized carbon are weakest. The order of bond strength is

$$sp > sp^2 > sp^3$$

This, too, is the order of the bond stiffness.

- The carbon–hydrogen stretching peaks of hydrogen atoms attached to *sp*-hybridized carbon atoms occur at highest frequencies, about 3300 cm^{-1}.

The carbon–hydrogen bond of a terminal alkyne (≡C—H) gives an absorption in the 3300-cm^{-1} region. We can see the absorption of the acetylenic (alkynyl) C—H bond of 1-heptyne at 3320 cm^{-1} in Fig. 2.13.

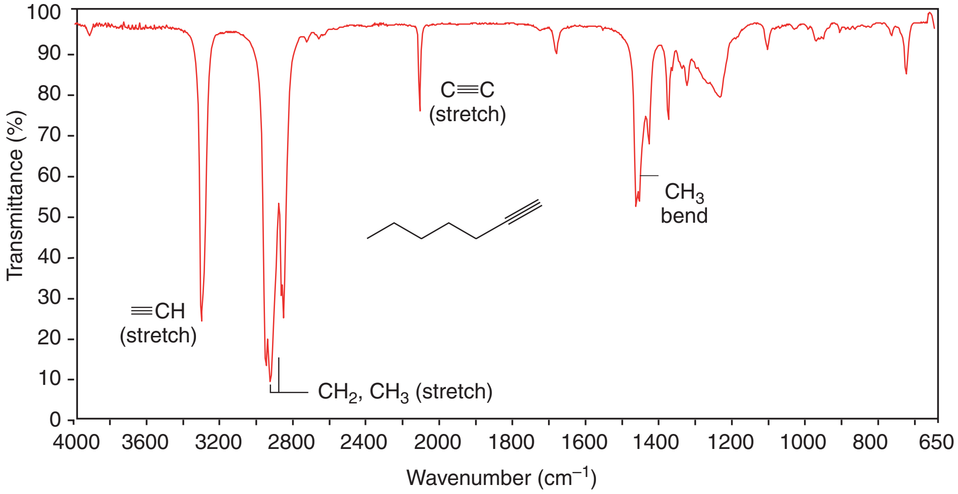


Figure 2.13 The IR spectrum of 1-heptyne.

- The carbon–hydrogen stretching peaks of hydrogen atoms attached to sp^2-hybridized carbon atoms occur in the 3000–3100-cm^{-1} region.

Thus, alkenyl C—H bonds and the C—H groups of aromatic rings give absorption peaks in this region. We can see the alkenyl C—H absorption peak at 3080 cm^{-1} in the spectrum of 1-octene (Fig. 2.14), and we can see the C—H absorption of the aromatic hydrogen atoms at 3090 cm^{-1} in the spectrum of methylbenzene (Fig. 2.12).

- The carbon–hydrogen stretching bands of hydrogen atoms attached to sp^3-hybridized carbon atoms occur at lowest frequencies, in the 2800–3000-cm^{-1} region.

We can see methyl and methylene absorption peaks in the spectra of octane (Fig. 2.11), methylbenzene (Fig. 2.12), 1-heptyne (Fig. 2.13), and 1-octene (Fig. 2.14).

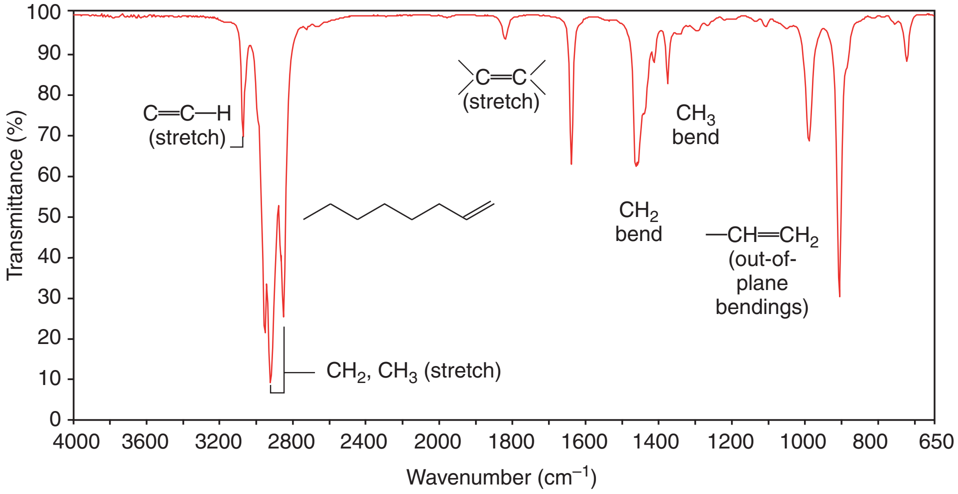


Figure 2.14 The IR spectrum of 1-octene.

Hydrocarbons also give absorption peaks in their IR spectra that result from carbon–carbon bond stretchings. Carbon–carbon single bonds normally give rise to very weak peaks that are usually of little use in assigning structures. More useful peaks arise from carbon–carbon multiple bonds, however.

- Carbon–carbon double bonds give absorption peaks in the 1620–1680-cm^{-1} region, and carbon–carbon triple bonds give absorption peaks between 2100 and 2260 cm^{-1}.

These absorptions are not usually strong ones, and they are absent if the double or triple bond is symmetrically substituted. (No dipole moment change will be associated with the vibration.) The stretchings of the carbon–carbon bonds of benzene rings usually give a set of characteristic sharp peaks in the 1450–1600-cm^{-1} region.

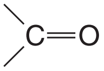
- Absorptions arising from carbon–hydrogen bending vibrations of alkenes occur in the 600–1000-cm^{-1} region. With the aid of a spectroscopy handbook, the exact location of these peaks can often be used as evidence for the ***substitution pattern of the double bond and its configuration***.

2.16B IR Spectra of Some Functional Groups Containing Heteroatoms

Helpful Hint

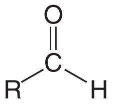
IR spectroscopy is an exceedingly useful tool for detecting functional groups.

Infrared spectroscopy gives us an invaluable method for recognizing quickly and simply the presence of certain functional groups in a molecule.

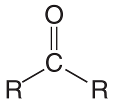
Carbonyl Functional Groups One important functional group that gives a prominent absorption peak in IR spectra is the **carbonyl group**,  . This group is present in aldehydes, ketones, esters, carboxylic acids, amides, and others.

- The carbon–oxygen double-bond stretching frequency of carbonyl groups gives a strong peak between 1630 and 1780 cm^{-1}.

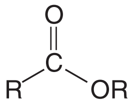
The exact location of the absorption depends on whether it arises from an aldehyde, ketone, ester, and so forth.



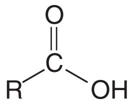
Aldehyde
1690–1740 cm^{-1}



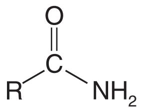
Ketone
1680–1750 cm^{-1}



Ester
1735–1750 cm^{-1}



Carboxylic acid
1710–1780 cm^{-1}

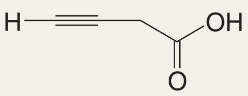


Amide
1630–1690 cm^{-1}

Solved Problem 2.7

A compound with the molecular formula $C_4H_4O_2$ has a strong sharp absorbance near 3300 cm^{-1}, absorbances in the 2800–3000-cm^{-1} region, and a sharp absorbance peak near 2200 cm^{-1}. It also has a strong broad absorbance in the 2500–3600-cm^{-1} region and a strong peak in the 1710–1780-cm^{-1} region. Propose a possible structure for the compound.

STRATEGY AND ANSWER The sharp peak near 3300 cm^{-1} is likely to arise from the stretching of a hydrogen attached to the *sp*-hybridized carbon of a triple bond. The sharp peak near 2200 cm^{-1}, where the triple bond of an alkyne stretches, is consistent with this. The peaks in the 2800–3000-cm^{-1} region suggest stretchings of the C—H bonds of alkyl groups, either CH_2 or CH_3 groups. The strong, broad absorbance in the 2500–3600-cm^{-1} region suggests a hydroxyl group arising from a carboxylic acid. The strong peak around 1710–1780 cm^{-1} is consistent with this since it could arise from the carbonyl group of a carboxylic acid. Putting all this together with the molecular formula suggests the compound is as shown at the right.



Review Problem 2.28

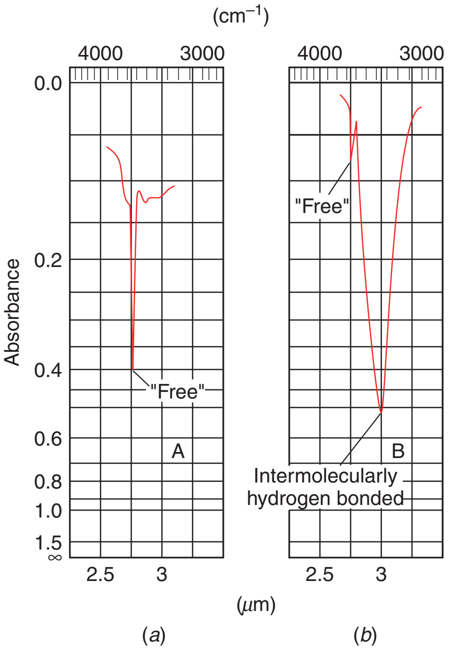
Use arguments based on resonance and electronegativity effects to explain the trend in carbonyl IR stretching frequencies from higher frequency for esters and carboxylic acids to lower frequencies for amides. (*Hint:* Use the range of carbonyl stretching frequencies for aldehydes and ketones as the "base" frequency range of an unsubstituted carbonyl group and consider the influence of electronegative atoms on the carbonyl group and/or atoms that alter the resonance hybrid of the carbonyl.) What does this suggest about the way the nitrogen atom influences the distribution of electrons in an amide carbonyl group?

Alcohols and Phenols

The **hydroxyl groups** of alcohols and phenols are also easy to recognize in IR spectra by their O—H stretching absorptions. These bonds also give us direct evidence for hydrogen bonding (Section 2.13B).

- The IR absorption of an alcohol or phenol O—H group is in the 3200–3550-cm^{-1} range, and most often it is broad.

The typical broadness of the peak is due to association of the molecules through hydrogen bonding (Section 2.13B), which causes a wider distribution of stretching frequencies for the O—H bond. If an alcohol or phenol is present as a very dilute solution in a solvent that cannot contribute to hydrogen bonding (e.g., CCl_4), O—H absorption occurs as a very sharp peak in the 3590–3650-cm^{-1} region. In very dilute solution in such a solvent or in the gas phase, formation of intermolecular hydrogen bonds does not take place because molecules of the analyte are too widely separated. A sharp peak in the 3590–3650-cm^{-1} region, therefore, is attributed to "free" (unassociated) hydroxyl groups. Increasing the concentration of the alcohol or phenol causes the sharp peak to be replaced by a broad band in the 3200–3550-cm^{-1} region. Hydroxyl absorptions in IR spectra of cyclohexylcarbinol (cyclohexylmethanol) run in dilute and concentrated solutions (Fig. 2.15) exemplify these effects.



(*a*) (*b*)

Figure 2.15 *(a)* The IR spectrum of an alcohol (cyclohexylcarbinol) in a dilute solution shows the sharp absorption of a "free" (non-hydrogen-bonded) hydroxyl group at 3600 cm^{-1}. *(b)* The IR spectrum of the same alcohol as a concentrated solution shows a broad hydroxyl group absorption at 3300 cm^{-1} due to hydrogen bonding. *(Reprinted with permission of John Wiley & Sons, Inc., from Silverstein, R., and Webster, F. X., Spectrometric Identification of Organic Compounds, Sixth Edition, p. 89. Copyright 1998.)*

Carboxylic Acids

The **carboxylic acid group** can also be detected by IR spectroscopy. If both carbonyl and hydroxyl stretching absorptions are present in an IR spectrum, there is good evidence for a carboxylic acid functional group (although it is possible that isolated carbonyl and hydroxyl groups could be present in the molecule).

- The hydroxyl absorption of a carboxylic acid is often very broad, extending from 3600 cm^{-1} to 2500 cm^{-1}.

Figure 2.16 shows the IR spectrum of propanoic acid.

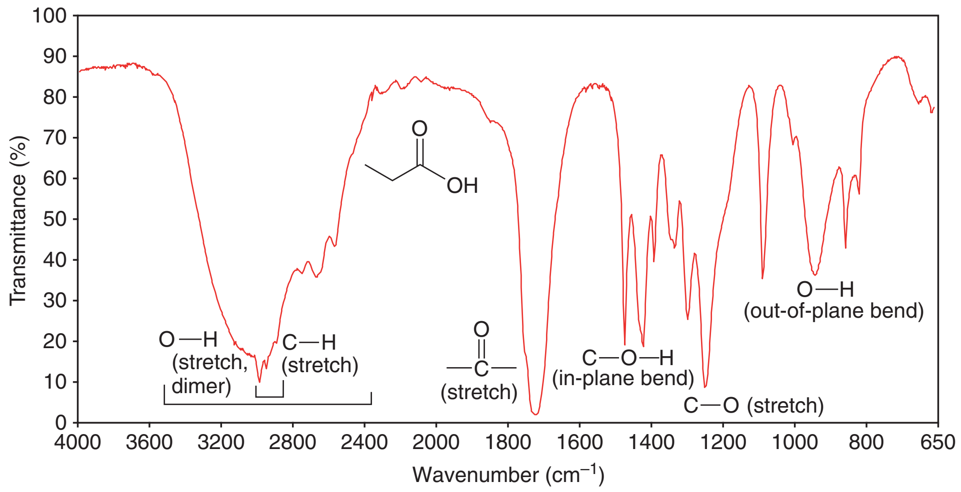


Figure 2.16 The IR spectrum of propanoic acid.

Amines IR spectroscopy also gives evidence for N—H bonds (see Figure 2.17).

- Primary (1°) and secondary (2°) amines give absorptions of moderate strength in the 3300–3500-cm^{-1} region.
- Primary amines exhibit two peaks in this region due to symmetric and asymmetric stretching of the two N—H bonds.
- Secondary amines exhibit a single peak.
- Tertiary amines show no N—H absorption because they have no such bond.
- A basic pH is evidence for any class of amines.

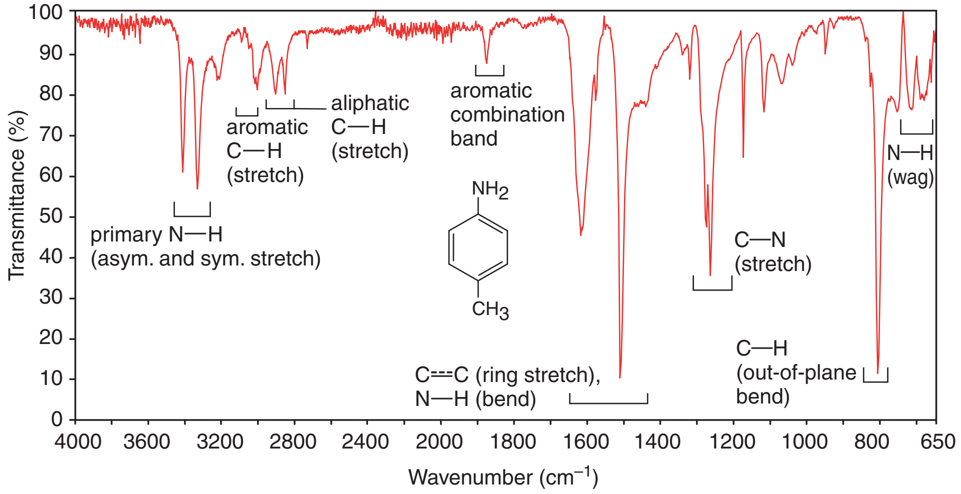
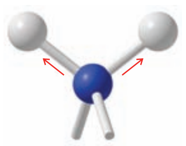
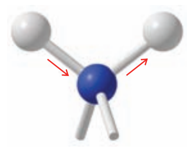


Figure 20.17 Annotated IR spectrum of 4-methylaniline.

RNH$_2$ (1° Amine)
Two peaks in
3300–3500-cm^{-1}
region

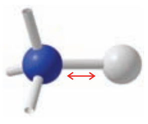


Symmetric stretching



Asymmetric stretching

R$_2$NH (2° Amine)
One peak in
3300–3500-cm^{-1}
region



Hydrogen bonding causes N—H stretching peaks of 1° and 2° amines to broaden. The NH groups of **amides** give similar absorption peaks and include a carbonyl absorption as well.

2.17 *Applications of Basic Principles*

We now review how certain basic principles apply to phenomena that we have studied in this chaper.

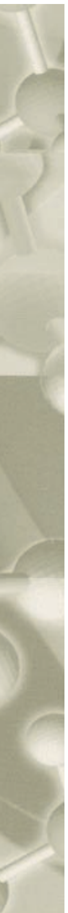
Polar Bonds Are Caused by Electronegativity Differences We saw in Section 2.2 that when atoms with different electronegativities are covalently bonded, the more electronegative atom will be negatively charged and the less electronegative atom will be positively charged. The bond will be a *polar bond* and it will have a *dipole moment*.

Dipole moments are important in explaining physical properties of molecules (as we shall review below), and in explaining infrared spectra. For a vibration to occur with the absorption of IR energy, the dipole moment of the molecule must change during the course of the vibration.

Opposite Charges Attract This principle underlies a map of electrostatic potential (MEP) (Section 2.2A). MEPs are generated on the basis of quantum mechanical calculations that involve moving an imaginary positive charge over the electron density surface of a molecule. If there is a strong attraction between the positive charge and the electron density surface, that region is colored red because it is most negative. Regions that are less negative are shaded green to yellow. Regions that are the least negative (or most positive) are colored blue.

This same principle is central to understanding physical properties of organic compounds (Section 2.13). All of the forces that operate between individual molecules (and thereby affect boiling points, melting points, and solubilities) are between oppositely charged molecules (ions) or between oppositely charged portions of molecules. Examples are ion–ion forces (Section 2.13A) that exist between oppositely charged ions in crystals of ionic compounds, dipole–dipole forces (Section 2.13B) that exist between oppositely charged portions of polar molecules and that include the very strong dipole–dipole forces that we call *hydrogen bonds*, and the weak *dispersion* or *London forces* that exist between portions of molecules that bear small temporary opposite charges.

Molecular Structure Determines Properties We learned in Section 2.13 how physical properties are related to molecular structure.



In This Chapter

In Chapter 2 you learned about families of organic molecules, some of their physical properties, and how we can use an instrumental technique called infrared spectroscopy to study them.

You learned that functional groups define the families to which organic compounds belong. At this point you should be able to name functional groups when you see them in structural formulas, and, when given the name of a functional group, draw a general example of its structure.

You also built on your knowledge of how electronegativity influences charge distribution in a molecule and how, together with three-dimensional structure, charge distribution influences the overall polarity of a molecule. Based on polarity and three-dimensional structure, you should be able to predict the kind and relative strength of electrostatic forces between molecules. With this understanding you will be able to roughly estimate physical properties such as melting point, boiling point, and solubility.

Lastly, you learned to use IR spectroscopy as an indicator of the family to which an organic compound belongs. IR spectroscopy provides signatures (in the form of spectra) that suggest which functional groups are present in a molecule.

If you know the concepts in Chapters 1 and 2 well, you will be on your way to having the solid foundation you need for success in organic chemistry. Keep up the good work (including your diligent homework habits)!

Key Terms and Concepts

WILEY PLUS

The key terms and concepts that are highlighted in **bold, blue text** within the chapter are defined in the glossary (at the back of the book) and have hyperlinked definitions in the accompanying ***WileyPLUS*** course (www.wileyplus.com).

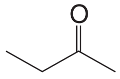
Problems

Note to Instructors: Many of the homework problems are available for assignment via *WileyPLUS*, an online teaching and learning solution.

FUNCTIONAL GROUPS AND STRUCTURAL FORMULAS

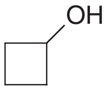
2.29 Classify each of the following compounds as an alkane, alkene, alkyne, alcohol, aldehyde, amine, and so forth.

(a)

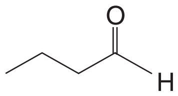


(b) $CH_3—C\equiv CH$

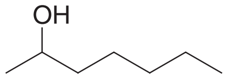
(c)



(d)



(e)



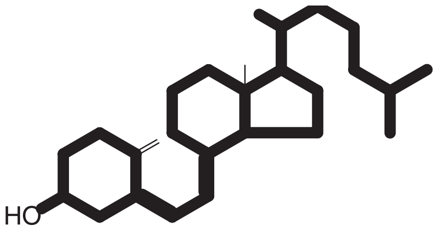
Obtained from oil of cloves

(f)

Sex attractant of the common housefly

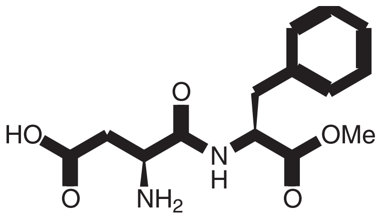
2.30 Identify all of the functional groups in each of the following compounds:

(a)



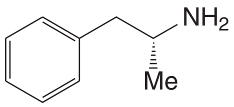
Vitamin D_3

(b)



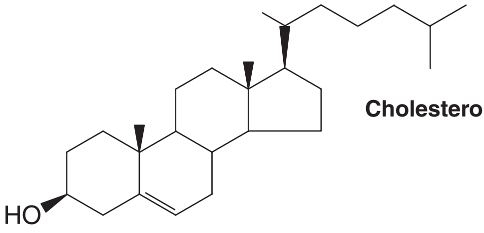
Aspartame

(c)



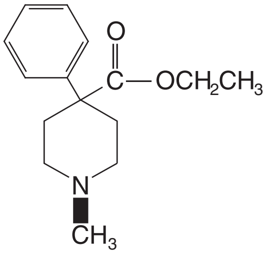
Amphetamine

(d)



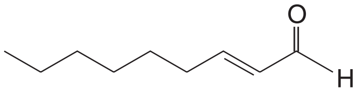
Cholesterol

(e)



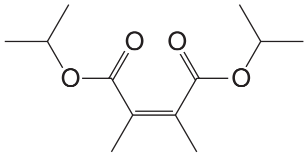
Demerol

(f)



A cockroach repellent found in cucumbers

(g)



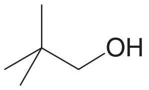
A synthetic cockroach repellent

2.31 There are four alkyl bromides with the formula C_4H_9Br. Write their structural formulas and classify each as to whether it is a primary, secondary, or tertiary alkyl bromide.

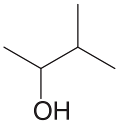
2.32 There are seven isomeric compounds with the formula $C_4H_{10}O$. Write their structures and classify each compound according to its functional group.

2.33 Classify the following alcohols as primary, secondary, or tertiary:

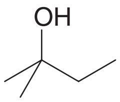
(a)



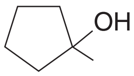
(b)



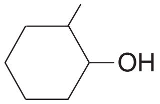
(c)



(d)

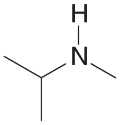


(e)

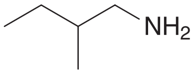


2.34 Classify the following amines as primary, secondary, or tertiary:

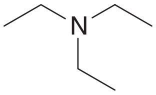
(a)



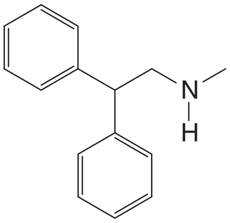
(b)



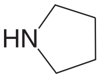
(c)



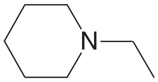
(d)



(e)



(f)

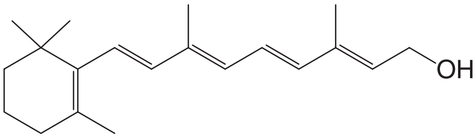


2.35 Write structural formulas for each of the following:

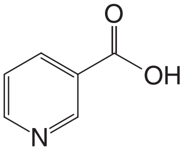
(a) Three ethers with the formula $C_4H_{10}O$.
(b) Three primary alcohols with the formula C_4H_8O.
(c) A secondary alcohol with the formula C_3H_6O.
(d) A tertiary alcohol with the formula C_4H_8O.
(e) Two esters with the formula $C_3H_6O_2$.
(f) Four primary alkyl halides with the formula $C_5H_{11}Br$.
(g) Three secondary alkyl halides with the formula $C_5H_{11}Br$.
(h) A tertiary alkyl halide with the formula $C_5H_{11}Br$.
(i) Three aldehydes with the formula $C_5H_{10}O$.
(j) Three ketones with the formula $C_5H_{10}O$.
(k) Two primary amines with the formula C_3H_9N.
(l) A secondary amine with the formula C_3H_9N.
(m) A tertiary amine with the formula C_3H_9N.
(n) Two amides with the formula C_2H_5NO.

PHYSICAL PROPERTIES

2.36 (**a**) Indicate the hydrophobic and hydrophilic parts of vitamin A and comment on whether you would expect it to be soluble in water. (**b**) Do the same for vitamin B_3 (also called niacin).



Vitamin A

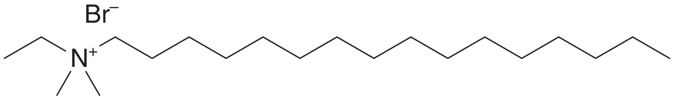


Vitamin B_3 or niacin

2.37 Hydrogen fluoride has a dipole moment of 1.83 D; its boiling point is 19.34°C. Ethyl fluoride (CH_3CH_2F) has an almost identical dipole moment and has a larger molecular weight, yet its boiling point is −37.7°C. Explain.

2.38 Why does one expect the cis isomer of an alkene to have a higher boiling point than the trans isomer?

2.39 Cetylethyldimethylammonium bromide is the common name for



, a compound with antiseptic properties. Predict its solubility behavior in water and in diethyl ether.

2.40 Which of the following solvents should be capable of dissolving ionic compounds?

(a) Liquid SO_2 **(b)** Liquid NH_3 **(c)** Benzene **(d)** CCl_4

2.41 Write a three-dimensional formula for each of the following molecules using the wedge–dashed wedge–line formalism. If the molecule has a net dipole moment, indicate its direction with an arrow, $+\!\!\!\longrightarrow$. If the molecule has no net dipole moment, you should so state. (You may ignore the small polarity of C—H bonds in working this and similar problems.)

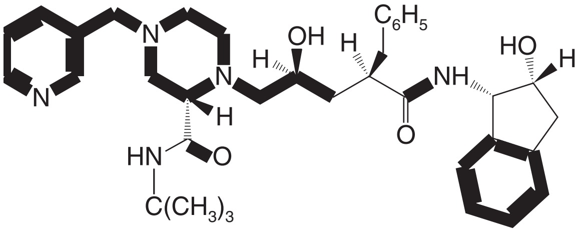
(a) CH_3F **(c)** CHF_3 **(e)** CH_2FCl **(g)** BeF_2 **(i)** CH_3OH
(b) CH_2F_2 **(d)** CF_4 **(f)** BCl_3 **(h)** CH_3OCH_3 **(j)** CH_2O

2.42 Consider each of the following molecules in turn: **(a)** dimethyl ether, $(CH_3)_2O$; **(b)** trimethylamine, $(CH_3)_3N$; **(c)** trimethylboron, $(CH_3)_3B$; and **(d)** dimethylberyllium, $(CH_3)_2Be$. Describe the hybridization state of the central atom (i.e., O, N, B, or Be) of each molecule, tell what bond angles you would expect at the central atom, and state whether the molecule would have a dipole moment.

2.43 Alkenes can interact with metal ions such as Ag^+. What is the nature of this interaction?

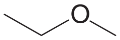
2.44 Analyze the statement: For a molecule to be polar, the presence of polar bonds is necessary, but it is not a sufficient requirement.

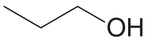
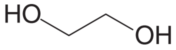
2.45 Identify all of the functional groups in Crixivan, an important drug in the treatment of AIDS.

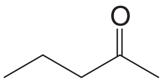
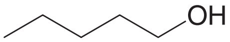


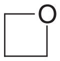
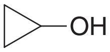
Crixivan (an HIV protease inhibitor)

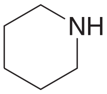
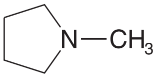
2.46 Which compound in each of the following pairs would have the higher boiling point? Explain your answers.

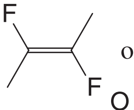
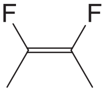
(a)  or 

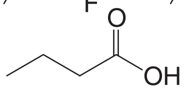
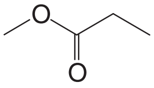
(b)  or 

(c)  or 

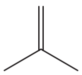
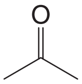
(d)  or 

(e)  or 

(f)  or 

(g)  or 

(h) Hexane, $CH_3(CH_2)_4CH_3$, or nonane, $CH_3(CH_2)_7CH_3$

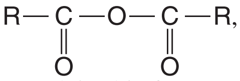
(i)  or 

IR SPECTROSCOPY

2.47 Predict the key IR absorption bands whose presence would allow each compound in pairs (**a**), (**c**), (**d**), (**e**), (**g**), and (**i**) from Problem 2.46 to be distinguished from each other.

2.48 The infrared spectrum of 1-hexyne exhibits a sharp absorption peak near 2100 cm^{-1} due to $C{\equiv}C$ stretching. However, 3-hexyne shows no absorption in that region. Explain.

2.49 The IR spectrum of propanoic acid (Fig. 2.17) indicates that the absorption for the O—H stretch of the carboxylic acid functional group is due to a hydrogen-bonded form. Draw the structure of two propanoic acid molecules showing how they could dimerize via hydrogen bonding.

2.50 In infrared spectra, the carbonyl group is usually indicated by a single strong and sharp absorption. However, in the case of carboxylic acid anhydrides,  $R-\underset{\underset{O}{\|}}{C}-O-\underset{\underset{O}{\|}}{C}-R$, two peaks are observed even though the two carbonyl groups are chemically equivalent. Explain this fact, considering what you know about the IR absorption of primary amines.

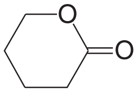
MULTICONCEPT PROBLEMS

2.51 Write structural formulas for four compounds with the formula C_3H_6O and classify each according to its functional group. Predict IR absorption frequencies for the functional groups you have drawn.

2.52 There are four amides with the formula C_3H_7NO. **(a)** Write their structures. **(b)** One of these amides has a melting and a boiling point that are substantially lower than those of the other three. Which amide is this? Explain your answer. **(c)** Explain how these amides could be differentiated on the basis of their IR spectra.

2.53 Write structures for all compounds with molecular formula C_4H_6O that would not be expected to exhibit infrared absorption in the 3200–3550-cm^{-1} and 1620–1780-cm^{-1} regions.

2.54 Cyclic compounds of the general type shown here are called lactones. What functional group does a lactone contain?



Challenge Problems

2.55 Two constitutional isomers having molecular formula C_4H_6O are both symmetrical in structure. In their infrared spectra, neither isomer when in dilute solution in CCl_4 (used because it is nonpolar) has absorption in the 3600-cm^{-1} region. Isomer A has absorption bands at approximately 3080, 1620, and 700 cm^{-1}. Isomer B has bands in the 2900-cm^{-1} region and at 1780 cm^{-1}. Propose a structure for A and two possible structures for B.

2.56 When two substituents are on the same side of a ring skeleton, they are said to be cis, and when on opposite sides, trans (analogous to use of those terms with 1,2-disubstituted alkene isomers). Consider stereoisomeric forms of 1,2-cyclopentanediol (compounds having a five-membered ring and hydroxyl groups on two adjacent carbons that are cis in one isomer and trans in the other). At high dilution in CCl_4, both isomers have an infrared absorption band at approximately 3626 cm^{-1} but only one isomer has a band at 3572 cm^{-1}. **(a)** Assume for now that the cyclopentane ring is coplanar (the interesting actuality will be studied later) and then draw and label the two isomers using the wedge–dashed wedge method of depicting the OH groups. **(b)** Designate which isomer will have the 3572-cm^{-1} band and explain its origin.

2.57 Compound C is asymmetric, has molecular formula $C_5H_{10}O$, and contains two methyl groups and a 3° functional group. It has a broad infrared absorption band in the 3200–3550-cm^{-1} region and no absorption in the 1620–1680-cm^{-1} region. **(a)** Propose a structure for C.

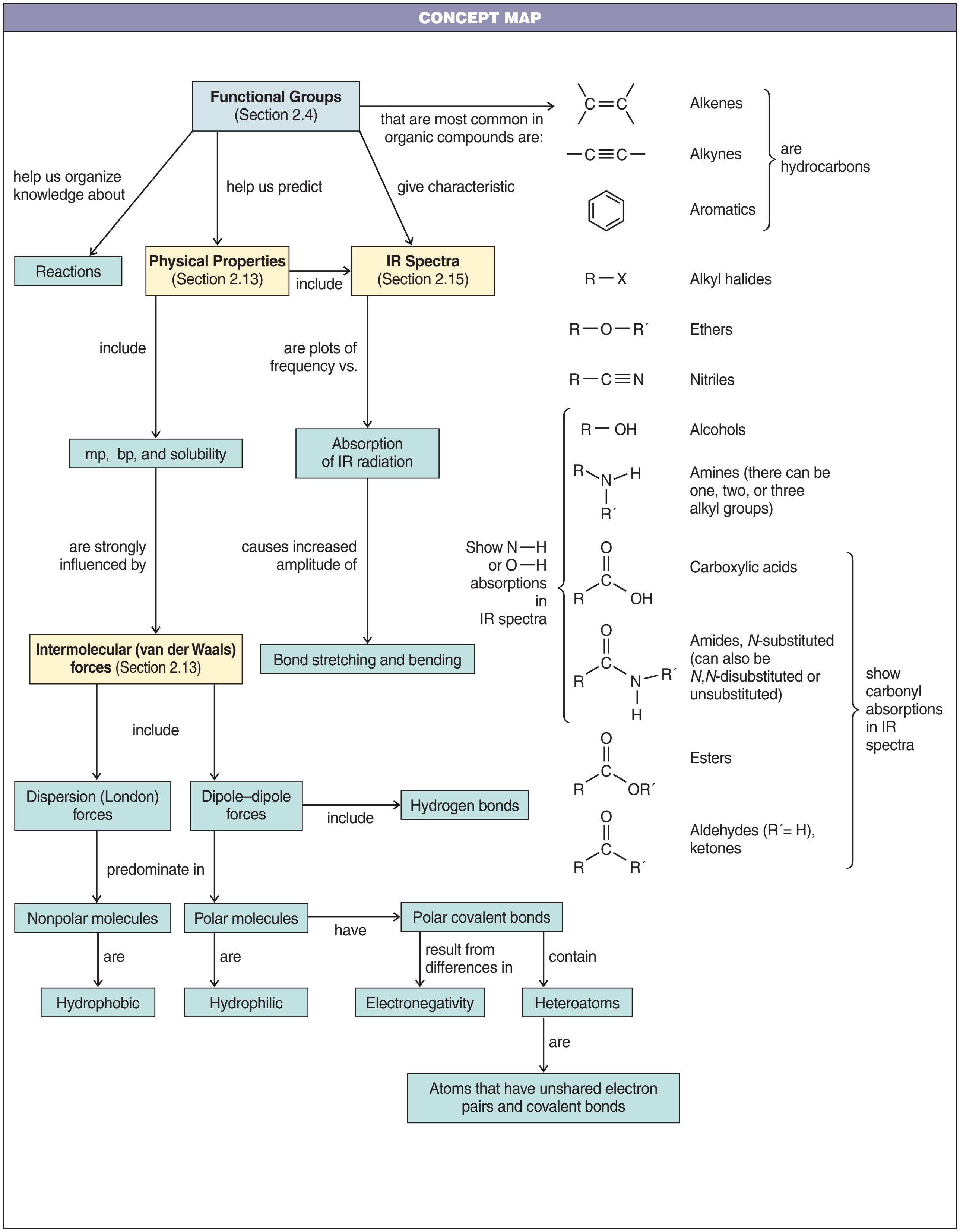
2.58 Examine the diagram showing an α-helical protein structure in section 2.13E. Between what specific atoms and of what functional groups are the hydrogen bonds formed that give the molecule its helical structure?

Learning Group Problems

Consider the molecular formula $C_4H_8O_2$.

1. Write structures for at least 15 different compounds that all have the molecular formula $C_4H_8O_2$ and contain functional groups presented in this chapter.
2. Provide at least one example each of a structure written using the dash format, the condensed format, the bond-line format, and the full three-dimensional format. Use your choice of format for the remaining structures.
3. Identify four different functional groups from among your structures. Circle and name them on the representative structures.
4. Predict approximate frequencies for IR absorptions that could be used to distinguish the four compounds representing these functional groups.
5. If any of the 15 structures you drew have atoms where the formal charge is other than zero, indicate the formal charge on the appropriate atom(s) and the overall charge for the molecule.
6. Identify which types of intermolecular forces would be possible in pure samples of all 15 compounds.
7. Pick five formulas you have drawn that represent a diversity of structures, and predict their order with respect to trend in increasing boiling point.
8. Explain your order of predicted boiling points on the basis of intermolecular forces and polarity.

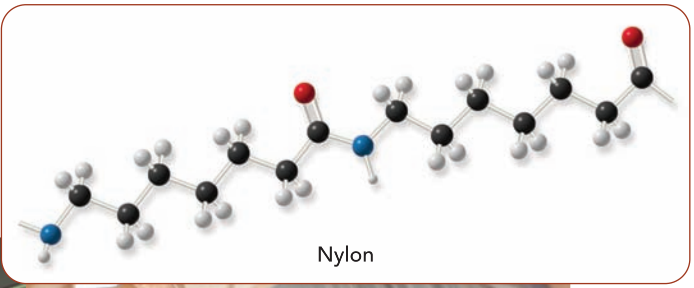
CONCEPT MAP


Functional Groups (Section 2.4)
help us organize knowledge about
Reactions
help us predict
Physical Properties (Section 2.13)
give characteristic
IR Spectra (Section 2.15)
include
that are most common in organic compounds are:
Alkenes
Alkynes
Aromatics
are hydrocarbons
R—X
Alkyl halides
R—O—R′
Ethers
R—C≡N
Nitriles
R—OH
Alcohols
Amines (there can be one, two, or three alkyl groups)
Carboxylic acids
Amides, N-substituted (can also be N,N-disubstituted or unsubstituted)
Show N—H or O—H absorptions in IR spectra
Esters
Aldehydes (R′= H), ketones
show carbonyl absorptions in IR spectra
include
mp, bp, and solubility
are plots of frequency vs.
Absorption of IR radiation
are strongly influenced by
causes increased amplitude of
Intermolecular (van der Waals) forces (Section 2.13)
Bond stretching and bending
include
Dispersion (London) forces
Dipole–dipole forces
include
Hydrogen bonds
predominate in
Nonpolar molecules
Polar molecules
have
Polar covalent bonds
are
are
result from differences in
contain
Hydrophobic
Hydrophilic
Electronegativity
Heteroatoms
are
Atoms that have unshared electron pairs and covalent bonds

3

An Introduction to Organic Reactions and Their Mechanisms

Acids and Bases


Nylon

To the uninitiated, a chemical reaction must seem like an act of magic. A chemist puts one or two reagents into a flask, waits for a time, and then takes from the flask one or more completely different compounds. It is, until we understand the details of the reaction, like a magician who puts apples and oranges in a hat, shakes it, and then pulls out rabbits and parakeets. We see a real-life example of this sort of "magic" in the photo above, where a student is shown pulling a strand of solid nylon from a flask that contains two immiscible solutions. This synthesis of nylon is not magic but it is indeed wonderful and amazing, and reactions like it have transformed our world.

One of our goals in this course will be, in fact, to try to understand how this chemical magic takes place. We will want to be able to explain *how the products of the reaction are formed*. This explanation will take the form of a **reaction mechanism**—**a description of the events that take place on a molecular level as reactants become products**. If, as is often the case, the reaction takes place in more than one step, we will want to know what chemical species, called **intermediates**, intervene between each step along the way.

By postulating a mechanism, we may take some of the magic out of the reaction, but we will put rationality in its place. Any mechanism we propose must be consistent with what we know about the reaction and with what we know about the reactivity of organic compounds generally. In later chapters we shall see how we can glean evidence for or against a given mechanism from studies of reaction rates, from isolating intermediates, and from spectroscopy. We cannot actually see the molecular events because molecules are too small, but from solid evidence and from good chemical intuition, we can propose reasonable mechanisms. If at some later time a valid experiment gives results that contradict our proposed mechanism, then we change it, because in the final analysis our mechanism must be consistent with all our experimental observations.

One of the most important things about approaching organic chemistry mechanistically is this: It helps us organize what otherwise might be an overwhelmingly complex body of knowledge into a form that makes it understandable. There are millions of organic compounds now known, and there are millions of reactions that these compounds undergo. If we had to learn them all by rote memorization, then we would soon give up. But, we don't have to do this. In the same way that functional groups help us organize compounds in a comprehensible way, mechanisms help us organize reactions. Fortunately, too, there is a relatively small number of basic mechanisms.

3.1 *Reactions and Their Mechanisms*

Virtually all organic reactions fall into one of four categories: *substitutions, additions, eliminations*, or *rearrangements*.

Substitutions are the characteristic reactions of saturated compounds such as alkanes and alkyl halides and of aromatic compounds (even though they are unsaturated). In a substitution, *one group replaces another*. For example, chloromethane reacts with sodium hydroxide to produce methyl alcohol and sodium chloride:

$$H_3C{-}Cl \quad + \quad Na^+OH^- \xrightarrow{H_2O} H_3C{-}OH \quad + \quad Na^+Cl^-$$

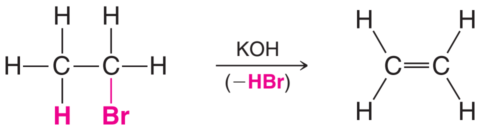
A substitution reaction

In this reaction a hydroxide ion from sodium hydroxide replaces the chlorine of methyl chloride. We shall study this reaction in detail in Chapter 6.

Additions are characteristic of compounds with multiple bonds. Ethene, for example, reacts with bromine by an addition. In an addition *all parts of the adding reagent appear in the product; two molecules become one*:

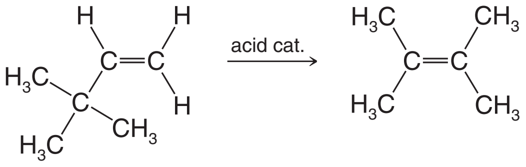
An addition reaction

Eliminations are the opposite of additions. *In an elimination one molecule loses the elements of another small molecule*. Elimination reactions give us a method for preparing compounds with double and triple bonds. In Chapter 7, for example, we shall study an important elimination called *dehydrohalogenation*, a reaction that is used to prepare alkenes. In dehydrohalogenation, as the word suggests, the elements of a hydrogen halide are eliminated. An alkyl halide becomes an alkene:



An elimination reaction

In a **rearrangement** *a molecule undergoes a reorganization of its constituent parts*. For example, heating the following alkene with a strong acid causes the formation of another isomeric alkene:



A rearrangement

In this rearrangement not only have the positions of the double bond and a hydrogen atom changed, but a methyl group has moved from one carbon to another.

In the following sections we shall begin to learn some of the principles that explain how these kinds of reactions take place.

3.1A Homolysis and Heterolysis of Covalent Bonds

Reactions of organic compounds always involve the making and breaking of covalent bonds. A covalent bond may break in two fundamentally different ways.

- When a bond breaks such that one fragment takes away both electrons of the bond, leaving the other fragment with an empty orbital, this kind of cleavage is called **heterolysis** (Greek: *hetero*, different, + *lysis*, loosening or cleavage). Heterolysis produces charged fragments or **ions** and is termed an **ionic reaction**. The broken bond is said to have cleaved *heterolytically*:

Helpful Hint

Notice in these illustrations that we have used curved arrows to show the movement of electrons. We will have more to say about this convention in Section 3.5, but for the moment notice that we use a double-barbed curved arrow to show the movement of a pair of electrons and a single-barbed curved arrow to show the movement of a single electron.

- When a bond breaks so that each fragment takes away one of the electrons of the bond, this process is called **homolysis** (Greek: *homo*, the same, + *lysis*). Homolysis produces fragments with unpaired electrons called **radicals**.

We shall postpone further discussions of reactions involving radicals and homolytic bond cleavage until we reach Chapter 10. At this point we focus our attention on reactions involving ions and heterolytic bond cleavage.

Heterolysis of a bond normally requires that the bond be polarized:

Polarization of a bond usually results from differing electronegativities (Section 2.2) of the atoms joined by the bond. The greater the difference in electronegativity, the greater is the polarization. In the instance just given, atom B is more electronegative than A.

Even with a highly polarized bond, heterolysis rarely occurs without assistance. The reason: *Heterolysis requires separation of oppositely charged ions*. Because oppositely charged ions attract each other, their separation requires considerable energy. Often, heterolysis is assisted by a molecule with an unshared pair that can form a bond to one of the atoms:

Formation of the new bond furnishes some of the energy required for the heterolysis.

3.2 Acid–Base Reactions

We begin our study of chemical reactions by examining some of the basic principles of acid–base chemistry. There are several reasons for doing this:

- Many of the reactions that occur in organic chemistry are either acid–base reactions themselves or they involve an acid–base reaction at some stage.
- Acid–base reactions are simple fundamental reactions that will enable you to see how chemists use curved arrows to represent mechanisms of reactions and how they depict the processes of bond breaking and bond making that occur as molecules react.

Acid–base reactions also allow us to examine important ideas about the relationship between the structures of molecules and their reactivity and to see how certain thermodynamic parameters can be used to predict how much of the product will be formed when a reaction reaches equilibrium. Acid–base reactions also provide an illustration of the important role solvents play in chemical reactions. They even give us a brief introduction to organic synthesis. Finally, acid–base chemistry is something that you will find familiar because of your studies in general chemistry. We begin, therefore, with a brief review.

3.2A Brønsted–Lowry Acids and Bases

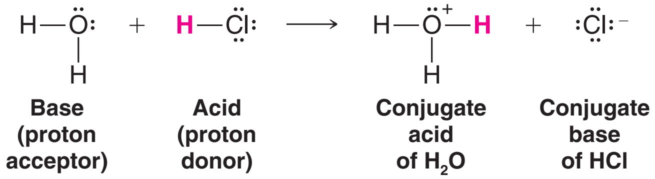
Two classes of acid–base reactions are fundamental in organic chemistry: Brønsted–Lowry and Lewis acid–base reactions. We start our discussion with Brønsted–Lowry acid–base reactions.

- Brønsted–Lowry acid–base reactions involve the transfer of protons.
- A **Brønsted–Lowry acid** is a substance that can donate (or lose) a proton.
- A **Brønsted–Lowry base** is a substance that can accept (or remove) a proton.

The color of hydrangea flowers depends in part on the relative acidity of their soil.

Let us consider some examples.

Hydrogen chloride (HCl), in its pure form, is a gas. When HCl gas is bubbled into water, the following reaction occurs.



In this reaction hydrogen chloride donates a proton; therefore it acts as a Brønsted–Lowry acid. Water accepts a proton from hydrogen chloride; thus water serves as a Brønsted–Lowry base. The products are a hydronium ion (H_3O^+) and chloride ion (Cl^-).

Just as we classified the reactants as either an acid or a base, we also classify the products in a specific way.

- The molecule or ion that forms when an acid loses its proton is called the **conjugate base** of that acid. In the above example, chloride ion is the conjugate base.
- The molecule or ion that forms when a base accepts a proton is called the **conjugate acid**. Hydronium ion is the conjugate acid of water.

Hydrogen chloride is considered a strong acid because transfer of its proton in water proceeds essentially to completion. Other strong acids that completely transfer a proton when dissolved in water are hydrogen iodide, hydrogen bromide, and sulfuric acid.

$$HI + H_2O \longrightarrow H_3O^+ + I^-$$
$$HBr + H_2O \longrightarrow H_3O^+ + Br^-$$
$$H_2SO_4 + H_2O \longrightarrow H_3O^+ + HSO_4^-$$
$$HSO_4^- + H_2O \rightleftharpoons H_3O^+ + SO_4^{2-}$$

Helpful Hint

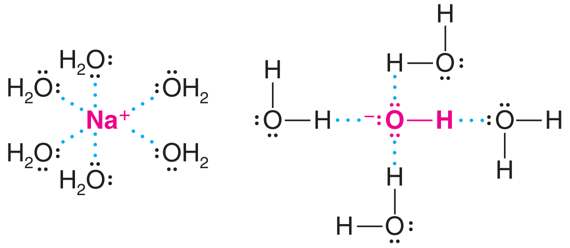
The extent to which an acid transfers protons to a base, such as water, is a measure of its strength as an acid. Acid strength is therefore a measure of the percentage of ionization and *not* of concentration.

Sulfuric acid is called a diprotic acid because it can transfer two protons. Transfer of the first proton occurs completely, while the second is transferred only to the extent of about 10% (hence the equilibrium arrows in the equation for the second proton transfer).

3.2B Acids and Bases in Water

- Hydronium ion is the strongest acid that can exist in water to any significant extent: Any stronger acid will simply transfer its proton to a water molecule to form hydronium ions.
- Hydroxide ion is the strongest base that can exist in water to any significant extent: Any base stronger than hydroxide will remove a proton from water to form hydroxide ions.

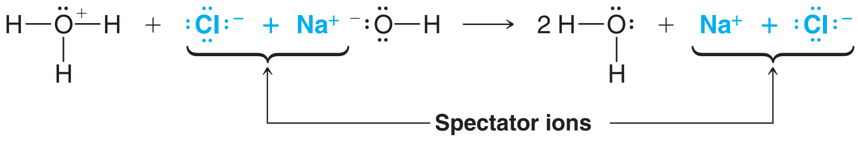
When an ionic compound dissolves in water the ions are solvated. With sodium hydroxide, for example, the positive sodium ions are stabilized by interaction with unshared electron pairs of water molecules, and the hydroxide ions are stabilized by hydrogen bonding of their unshared electron pairs with the partially positive hydrogens of water molecules.



Solvated sodium ion **Solvated hydroxide ion**

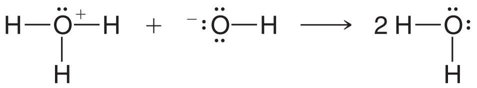
When an aqueous solution of sodium hydroxide is mixed with an aqueous solution of hydrogen chloride (hydrochloric acid), the reaction that occurs is between hydronium and hydroxide ions. The sodium and chloride ions are called **spectator ions** because they play no part in the acid–base reaction:

Total Ionic Reaction



$$H_3O^+ + \underbrace{Cl^- + Na^+}\ {}^-OH \longrightarrow 2\,H_2O + \underbrace{Na^+ + Cl^-}$$

Net Reaction



$$H_3O^+ + {}^-OH \longrightarrow 2\,H_2O$$

What we have just said about hydrochloric acid and aqueous sodium hydroxide is true when solutions of all aqueous strong acids and bases are mixed. The net ionic reaction is simply

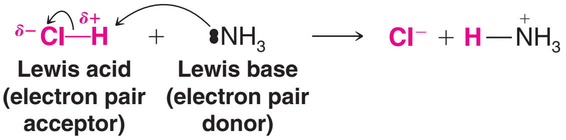
$$H_3O^+ + OH^- \longrightarrow 2\,H_2O$$

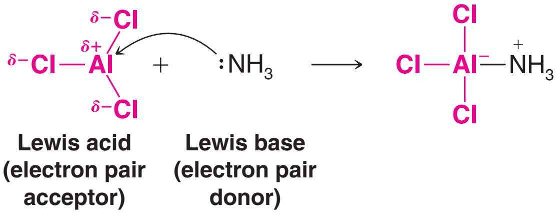
3.3 *Lewis Acids and Bases*

In 1923 G. N. Lewis proposed a theory that significantly broadened the understanding of acids and bases. As we go along we shall find that an understanding of **Lewis acid–base theory** is exceedingly helpful to understanding a variety of organic reactions. Lewis proposed the following definitions for acids and bases.

- Acids are electron pair acceptors.
- Bases are electron pair donors.

In Lewis acid–base theory, proton donors are not the only acids; many other species are acids as well. Aluminum chloride, for example, reacts with ammonia in the same way that a proton donor does. Using curved arrows to show the donation of the electron pair of ammonia (the Lewis base), we have the following examples:





Helpful Hint

Verify for yourself that you can calculate the formal charges in these structures.

In the reaction with hydrogen chloride above, notice that the electron pair acceptor (the proton) must also lose an electron pair as the new bond is formed with nitrogen. This is necessary because the hydrogen atom had a full valence shell of electrons at the start. On the other hand, because the valence shell of the aluminum atom in aluminum chloride was not full at the beginning (it had only a sextet of valence electrons), it can accept an electron pair without breaking any bonds. The aluminum atom actually achieves an octet by accepting the pair from nitrogen, although it gains a formal negative charge. When it accepts the electron pair, aluminum chloride is, in the Lewis definition, *acting as an acid*.

Bases are much the same in the Lewis theory and in the Brønsted–Lowry theory, because in the Brønsted–Lowry theory a base must donate a pair of electrons in order to accept a proton.

- The Lewis theory, by virtue of its broader definition of acids, allows acid–base theory to include all of the Brønsted–Lowry reactions and, as we shall see, a great many others. Most of the reactions we shall study in organic chemistry involve Lewis acid–base interactions, and a sound understanding of Lewis acid–base chemistry will help greatly.

Any *electron-deficient atom* can act as a Lewis acid. Many compounds containing group IIIA elements such as boron and aluminum are Lewis acids because group IIIA atoms have only a sextet of electrons in their outer shell. Many other compounds that have atoms with vacant orbitals also act as Lewis acids. Zinc and iron(III) halides (ferric halides) are frequently used as Lewis acids in organic reactions.

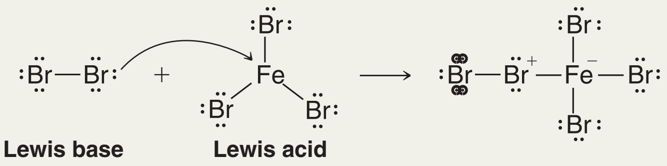
Carbonic anhydrase

A zinc ion acts as a Lewis acid in the mechanism of the enzyme carbonic anhydrase (Chapter 24).

Solved Problem 3.1

Write an equation that shows the Lewis acid and Lewis base in the reaction of bromine (Br_2) with ferric bromide ($FeBr_3$).

ANSWER



3.3A Opposite Charges Attract

- In Lewis acid–base theory, as in many organic reactions, the attraction of oppositely charged species is fundamental to reactivity.

As one further example, we consider boron trifluoride, an even more powerful Lewis acid than aluminum chloride, and its reaction with ammonia. The calculated structure for boron

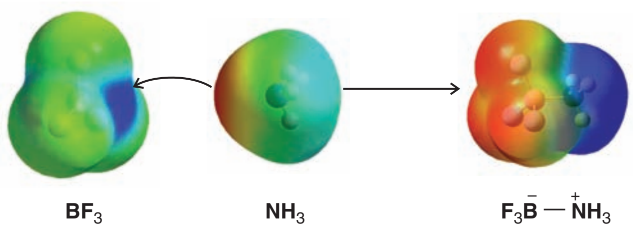


Figure 3.1 Electrostatic potential maps for BF_3 and NH_3 and the product that results from reaction between them. Attraction between the strongly positive region of BF_3 and the negative region of NH_3 causes them to react. The electrostatic potential map for the product shows that the fluorine atoms draw in the electron density of the formal negative charge, and the nitrogen atom, with its hydrogens, carries the formal positive charge.

trifluoride in Fig. 3.1 shows **electrostatic potential** at its van der Waals surface (like that in Section 2.2A for HCl). It is obvious from this figure (and you should be able to predict this) that BF_3 has substantial positive charge centered on the boron atom and negative charge located on the three fluorines. (The convention in these structures is that blue represents relatively positive areas and red represents relatively negative areas.) On the other hand, the surface electrostatic potential for ammonia shows (as you would expect) that substantial negative charge is localized in the region of ammonia's nonbonding electron pair. Thus, the electrostatic properties of these two molecules are perfectly suited for a Lewis acid–base reaction. When the expected reaction occurs between them, the nonbonding electron pair of ammonia attacks the boron atom of boron trifluoride, filling boron's valence shell. The boron now carries a formal negative charge and the nitrogen carries a formal positive charge. This separation of charge is borne out in the electrostatic potential map for the product shown in Fig. 3.1. Notice that substantial negative charge resides in the BF_3 part of the molecule, and substantial positive charge is localized near the nitrogen.

Although calculated electrostatic potential maps like these illustrate charge distribution and molecular shape well, it is important that you are able to draw the same conclusions based on what you would have predicted about the structures of BF_3 and NH_3 and their reaction product using orbital hybridization (Sections 1.12–1.14), VSEPR models (Section 1.16), consideration of formal charges (Section 1.7), and electronegativity (Sections 1.4A and 2.2).

Helpful Hint

The need for a firm understanding of structure, formal charges, and electronegativity can hardly be emphasized enough as you build a foundation of knowledge for learning organic chemistry.

Review Problem 3.1

Write equations showing the Lewis acid–base reaction that takes place when:

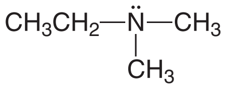
(a) Methanol (CH_3OH) reacts with BF_3.

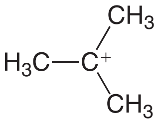
(b) Chloromethane (CH_3Cl) reacts with $AlCl_3$.

(c) Dimethyl ether (CH_3OCH_3) reacts with BF_3.

Review Problem 3.2

Which of the following are potential Lewis acids and which are potential Lewis bases?

(a)

$CH_3CH_2—\ddot{N}(CH_3)—CH_3$

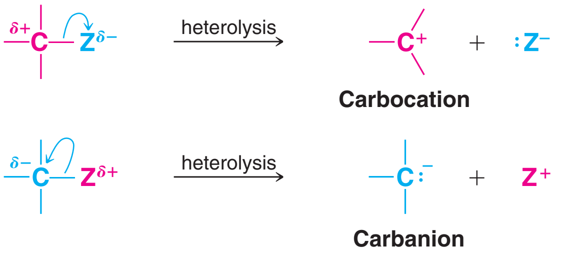
(b)

$H_3C—\overset{+}{C}(CH_3)_2$

(c) $(C_6H_5)_3P:$

(d) $:\ddot{\underset{..}{Br}}:^-$ **(e)** $(CH_3)_3B$ **(f)** $H:^-$

3.4 *Heterolysis of Bonds to Carbon: Carbocations and Carbanions*

Heterolysis of a bond to a carbon atom can lead to either of two ions: either to an ion with a positive charge on the carbon atom, called a **carbocation**, or to an ion with a negatively charged carbon atom, called a **carbanion**:

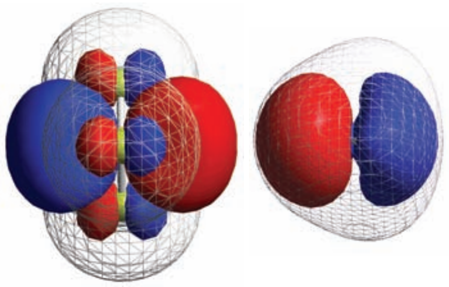


THE CHEMISTRY OF . . .

HOMOs and LUMOs in Reactions

The calculated lowest unoccupied molecular orbital (**LUMO**) for BF_3 is shown by solid red and blue lobes. Most of the volume represented by the LUMO corresponds to the empty *p* orbital in the *sp*^{2}-hybridized state of BF_3 (located perpendicular to the plane of the atoms). This orbital is where electron density fills (bonding occurs) when BF_3 is attacked by NH_3. The van der Waals surface electron density of BF_3 is indicated by the mesh. As the structure shows, the LUMO extends beyond the electron density surface, and hence it is easily accessible for reaction.

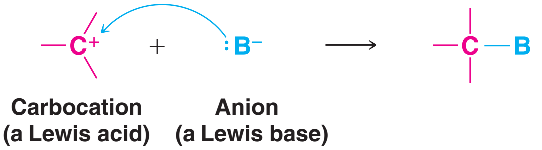
The highest occupied molecular orbital (**HOMO**) of ammonia, where the nonbonding pair resides, is shown by red and blue lobes in its structure. When the reaction occurs, the electron density from the HOMO of ammonia is transferred to the LUMO of boron trifluoride. This interaction involving the HOMO of one molecule with the LUMO of another is, from a molecular orbital perspective, the way reactions occur.

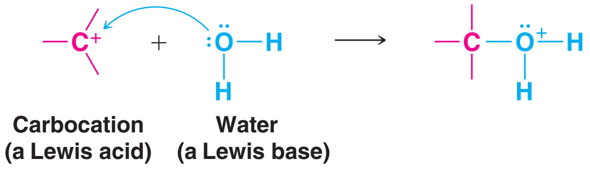


The **LUMO** of BF_3 (left) and the **HOMO** of NH_3 (right).

- Carbocations are electron deficient. They have only six electrons in their valence shell, and because of this, carbocations are Lewis acids.

In this way they are like BF_3 and $AlCl_3$. Most carbocations are also short-lived and highly reactive. They occur as intermediates in some organic reactions. Carbocations react rapidly with Lewis bases—with molecules or ions that can donate the electron pair that they need to achieve a stable octet of electrons (i.e., the electronic configuration of a noble gas):



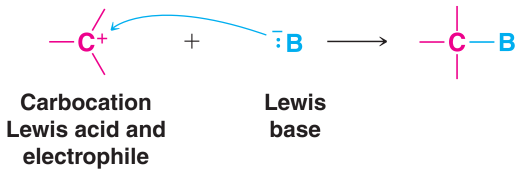


- **Carbanions** are electron rich. They are anions and have an unshared electron pair. Carbanions, therefore, are **Lewis bases and react accordingly** (Section 3.3).

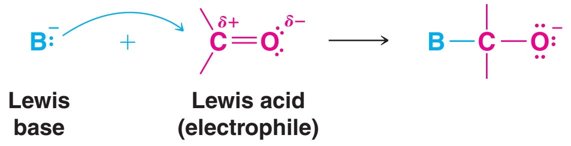
3.4A Electrophiles and Nucleophiles

Because carbocations are electron-seeking reagents chemists call them **electrophiles** (meaning electron-loving).

- **Electrophiles are reagents that seek electrons so as to achieve a stable shell of electrons like that of a noble gas.**
- **All Lewis acids are electrophiles**. By accepting an electron pair from a Lewis base, a carbocation fills its valence shell.



- **Carbon atoms that are electron poor because of bond polarity, but are not carbocations, can also be electrophiles**. They can react with the electron-rich centers of Lewis bases in reactions such as the following:

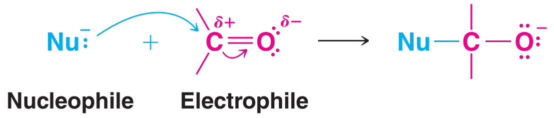


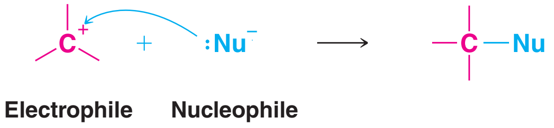
Carbanions are Lewis bases. Carbanions seek a proton or some other positive center to which they can donate their electron pair and thereby neutralize their negative charge.

When a Lewis base *seeks a positive center other than a proton, especially that of a carbon atom*, chemists call it a **nucleophile** (meaning nucleus loving; the *nucleo-* part of the name comes from *nucleus*, the positive center of an atom).

- A **nucleophile** **is a Lewis base that seeks a positive center such as a positively charged carbon atom**.

Since electrophiles are also Lewis acids (electron pair acceptors) and nucleophiles are Lewis bases (electron pair donors), why do chemists have two terms for them? The answer is that *Lewis acid* and *Lewis base* are terms that are used generally, but when one or the other reacts to form a bond to a carbon atom, we usually call it an *electrophile* or a *nucleophile*.





3.5 How to *Use Curved Arrows in Illustrating Reactions*

Up to this point we have not indicated how bonding changes occur in the reactions we have presented, but this can easily be done using curved-arrow notation.

Curved arrows

- show the direction of electron flow in a reaction mechanism.
- point from the source of an electron pair to the atom receiving the pair. (Curved arrows can also show the movement of single electrons. We shall discuss reactions of this type in a later chapter.)
- always show the flow of electrons from a site of higher electron density to a site of lower electron density.
- **never** show the movement of atoms. Atoms are assumed to follow the flow of the electrons.

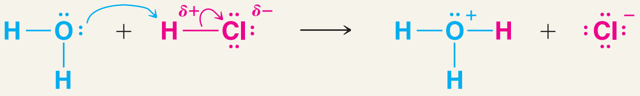
The reaction of hydrogen chloride with water provides a simple example of how to use curved arrow notation. Here we invoke the first of many "A Mechanism for the Reaction" boxes, in which we show every key step in a mechanism using color-coded formulas accompanied by explanatory captions.

A MECHANISM FOR THE REACTION

Reaction of Water with Hydrogen Chloride: The Use of Curved Arrows

REACTION $H_2O + HCl \longrightarrow H_3O^+ + Cl^-$

MECHANISM



A water molecule uses one of the nonbonding electron pairs to form a bond to a proton of HCl. The bond between the hydrogen and chlorine breaks, and the electron pair goes to the chlorine atom.

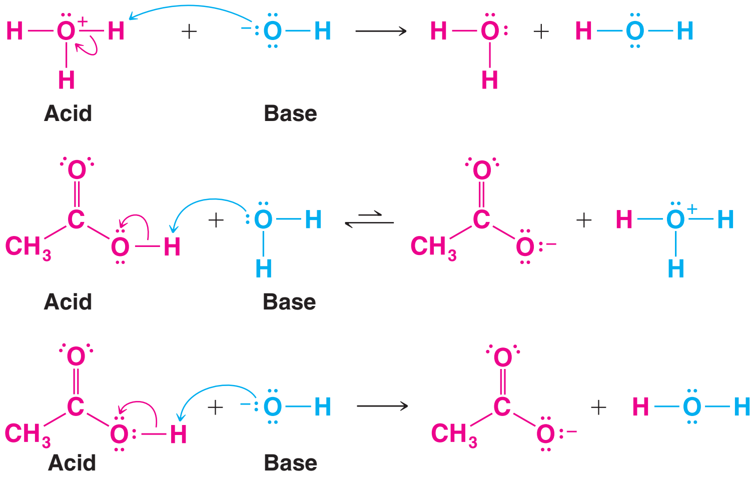
This leads to the formation of a hydronium ion and a chloride ion.

Helpful Hint

Curved arrows point *from* electrons *to* the atom receiving the electrons.

The curved arrow begins with a covalent bond or unshared electron pair (a site of higher electron density) and points toward a site of electron deficiency. We see here that as the water molecule collides with a hydrogen chloride molecule, it uses one of its unshared electron pairs (shown in blue) to form a bond to the proton of HCl. This bond forms because the negatively charged electrons of the oxygen atom are attracted to the positively charged proton. As the bond between the oxygen and the proton forms, the hydrogen–chlorine bond of HCl breaks, and the chlorine of HCl departs with the electron pair that formerly bonded it to the proton. (If this did not happen, the proton would end up forming two covalent bonds, which, of course, a proton cannot do.) We, therefore, use a curved arrow to show the bond cleavage as well. By pointing from the bond to the chlorine, the arrow indicates that the bond breaks and the electron pair leaves with the chloride ion.

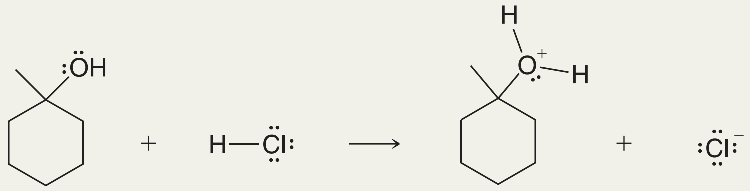
The following acid–base reactions give other examples of the use of the curved-arrow notation:



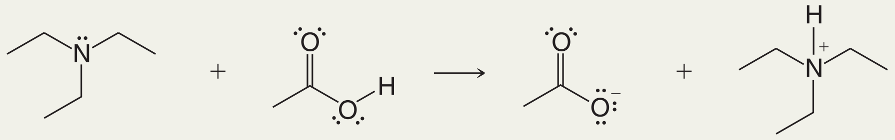
Solved Problem 3.2

Add curved arrows to the following reactions to indicate the flow of electrons for all of the bond-forming and bond-breaking steps.

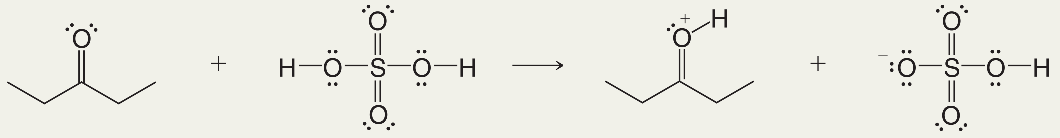
(a)



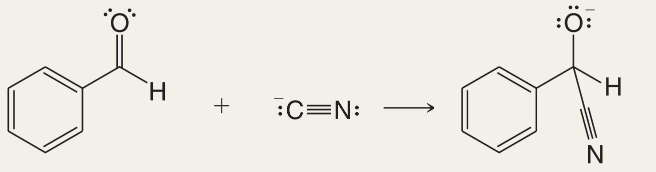
(b)



(c)



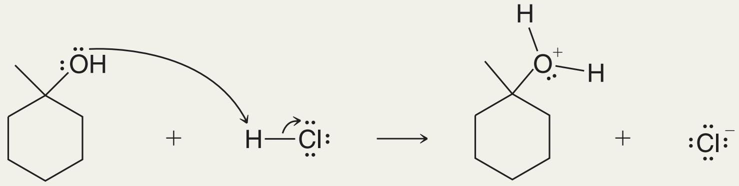
(d)



STRATEGY AND ANSWER Recall the rules for use of curved arrows presented at the beginning of Section 3.5. Curved arrows point from the source of an electron pair to the atom receiving the pair, and always point from a site of higher electron density to a site of lower electron density. We must also not exceed two electrons for a hydrogen atom, or an octet of electrons for any elements in the second row of the periodic table. We must also account for the formal charges on atoms and write equations whose charges are balanced.

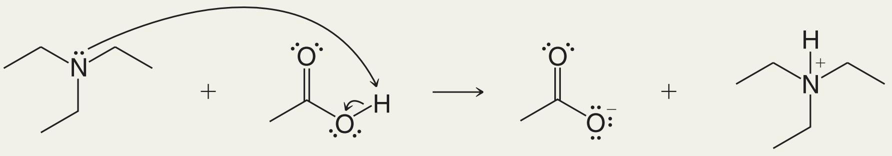
In (a), the hydrogen atom of HCl is partially positive (electrophilic) due to the electronegativity of the chlorine atom. The alcohol oxygen is a source of electrons (a Lewis base) that can be given to this partially positive proton. The proton must lose a pair of electrons as it gains a pair, however, and thus the chloride ion accepts a pair of electrons from the bond it had with the hydrogen atom as the hydrogen becomes bonded to the alcohol oxygen.

(a)



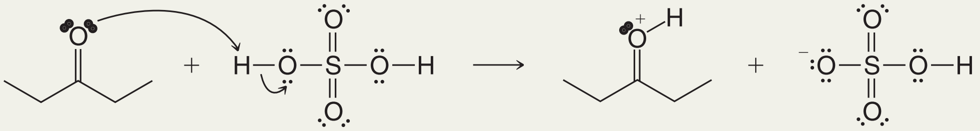
In (b), the carboxylic acid hydrogen is partially positive and therefore electrophilic, and the amine provides an unshared pair of electrons that forms a bond with the carboxylic acid hydrogen, causing departure of a carboxylate anion.

(b)



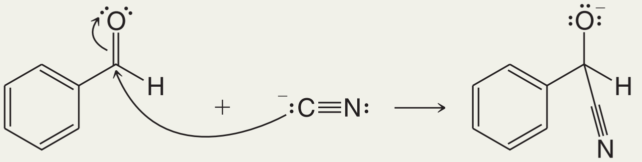
The circumstances in (c) are quite similar to (a) and (b). In this case the electrophile is a proton of sulfuric acid.

(c)



In (d), the aldehyde carbon is electrophilic due to the electronegativity of the carbonyl oxygen. The cyanide anion acts as a Lewis base, donating an electron pair to the carbonyl carbon, and causing an electron pair to shift to the oxygen so that no atom has more than an octet of electrons.

(d)

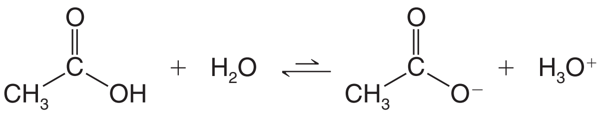


Review Problem 3.3

Use the curved-arrow notation to write the reaction that would take place between dimethylamine $(CH_3)_2NH$ and boron trifluoride. Identify the Lewis acid and Lewis base and assign appropriate formal charges.

3.6 The Strength of Brønsted–Lowry Acids and Bases: K_a and pK_a

In contrast to the strong acids, such as HCl and H_2SO_4, acetic acid is a much weaker acid. When acetic acid dissolves in water, the following reaction does not proceed to completion:



Experiments show that in a 0.1*M* solution of acetic acid at 25°C only about 1% of the acetic acid molecules ionize by transferring their protons to water. Therefore, acetic acid is a weak acid. As we shall see next, **acid strength** is characterized in terms of acidity constant (K_a) or pK_a values.

3.6A The Acidity Constant, K_a

Because the reaction that occurs in an aqueous solution of acetic acid is an equilibrium, we can describe it with an expression for the **equilibrium constant (K_{eq})**:

$$K_{eq} = \frac{[H_3O^+][CH_3CO_2^-]}{[CH_3CO_2H][H_2O]}$$

For dilute aqueous solutions, the concentration of water is essentially constant (~55.5*M*), so we can rewrite the expression for the equilibrium constant in terms of a new constant **(K_a)** called the **acidity constant**:

$$K_a = K_{eq}[H_2O] = \frac{[H_3O^+][CH_3CO_2^-]}{[CH_3CO_2H]}$$

At 25°C, the acidity constant for acetic acid is 1.76×10^{-5}.

We can write similar expressions for any weak acid dissolved in water. Using a generalized hypothetical acid (HA), the reaction in water is

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

and the expression for the acidity constant is

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

Because the concentrations of the products of the reaction are written in the numerator and the concentration of the undissociated acid in the denominator, **a large value of K_a means the acid is a strong acid and a small value of K_a means the acid is a weak acid**. If the K_a is greater than 10, the acid will be, for all practical purposes, completely dissociated in water at concentrations less than 0.01M.

Review Problem 3.4

Formic acid (HCO_2H) has $K_a = 1.77 \times 10^{-4}$. **(a)** What are the molar concentrations of the hydronium ion and formate ion (HCO_2^-) in a 0.1M aqueous solution of formic acid? **(b)** What percentage of the formic acid is ionized?

3.6B Acidity and pK_a

Chemists usually express the acidity constant, K_a, as its negative logarithm, **pK_a**:

$$pK_a = -\log K_a$$

This is analogous to expressing the hydronium ion concentration as pH:

$$pH = -\log[H_3O^+]$$

For acetic acid the pK_a is 4.75:

$$pK_a = -\log(1.76 \times 10^{-5}) = -(-4.75) = 4.75$$

Helpful Hint

K_a and pK_a are indicators of acid strength.

Notice that there is an inverse relationship between the magnitude of the pK_a and the strength of the acid.

- **The larger the value of the pK_a, the weaker is the acid**.

For example, acetic acid with pK_a = 4.75 is a weaker acid than trifluoroacetic acid with pK_a = 0 (K_a = 1). Hydrochloric acid with pK_a = −7 ($K_a = 10^7$) is a far stronger acid than trifluoroacetic acid. (It is understood that a positive pK_a is larger than a negative pK_a.)

$$CH_3CO_2H < CF_3CO_2H < HCl$$

$$pK_a = 4.75 \quad pK_a = 0 \quad pK_a = -7$$

Weak acid **Very strong acid**

Increasing acid strength

Table 3.1 lists pK_a values for a selection of acids relative to water as the base. The values in the middle pK_a range of the table are the most accurate because they can be measured in aqueous solution. Special methods must be used to estimate the pK_a values for the very strong acids at the top of the table and for the very weak acids at the bottom.* The pK_a values for these very strong and weak acids are therefore approximate. All of the acids that we shall consider in this book will have strengths in between that of ethane (an extremely weak acid) and that of $HSbF_6$ (an acid that is so strong that it is called a "superacid"). As you examine Table 3.1, take care not to lose sight of the vast range of acidities that it represents (a factor of 10^{62}).

---

*Acids that are stronger than a hydronium ion and bases that are stronger than a hydroxide ion react completely with water (a phenomenon called the **leveling effect**; see Sections 3.2B and 3.15). Therefore, it is not possible to measure acidity constants for these acids in water. Other solvents and special techniques are used, but we do not have the space to describe those methods here.

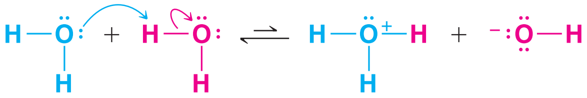
TABLE 3.1 Relative Strength of Selected Acids and Their Conjugate Bases

	Acid	Approximate pK_a	Conjugate Base	
Strongest acid	$HSbF_6$	<-12	SbF_6^-	Weakest base
	HI	-10	I^-	
	H_2SO_4	-9	HSO_4^-	
	HBr	-9	Br^-	
	HCl	-7	Cl^-	
	$C_6H_5SO_3H$	-6.5	$C_6H_5SO_3^-$	
	$(CH_3)_2\overset{+}{O}H$	-3.8	$(CH_3)_2O$	
	$(CH_3)_2C{=}\overset{+}{O}H$	-2.9	$(CH_3)_2C{=}O$	
	$CH_3\overset{+}{O}H_2$	-2.5	CH_3OH	
	H_3O^+	-1.74	H_2O	
	HNO_3	-1.4	NO_3^-	
	CF_3CO_2H	0.18	$CF_3CO_2^-$	
	HF	3.2	F^-	
	$C_6H_5CO_2H$	4.21	$C_6H_5CO_2^-$	Increasing base strength
	$C_6H_5NH_3^+$	4.63	$C_6H_5NH_2$	
Increasing acid strength	CH_3CO_2H	4.75	$CH_3CO_2^-$	
	H_2CO_3	6.35	HCO_3^-	
	$CH_3COCH_2COCH_3$	9.0	$CH_3COHCOCH_3$	
	NH_4^+	9.2	NH_3	
	C_6H_5OH	9.9	$C_6H_5O^-$	
	HCO_3^-	10.2	CO_3^{2-}	
	$CH_3NH_3^+$	10.6	CH_3NH_2	
	H_2O	15.7	OH^-	
	CH_3CH_2OH	16	$CH_3CH_2O^-$	
	$(CH_3)_3COH$	18	$(CH_3)_3CO^-$	
	CH_3COCH_3	19.2	$^-CH_2COCH_3$	
	$HC{\equiv}CH$	25	$HC{\equiv}C^-$	
	H_2	35	H^-	
	NH_3	38	NH_2^-	
	$CH_2{=}CH_2$	44	$CH_2{=}CH^-$	
Weakest acid	CH_3CH_3	50	$CH_3CH_2^-$	Strongest base

Review Problem 3.5

(a) An acid (HA) has $K_a = 10^{-7}$. What is its pK_a? **(b)** Another acid (HB) has $K_a = 5$; what is its pK_a? **(c)** Which is the stronger acid?

Water, itself, is a very weak acid and undergoes self-ionization even in the absence of acids and bases:



In pure water at 25°C, the concentrations of hydronium and hydroxide ions are equal to $10^{-7}M$. Since the concentration of water in pure water is 55.5M, we can calculate the K_a for water.

$$K_a = \frac{[H_3O^+][OH^-]}{[H_2O]} \qquad K_a = \frac{(10^{-7})(10^{-7})}{55.5} = 1.8 \times 10^{-16} \qquad pK_a = 15.7$$

Review Problem 3.6

Show calculations proving that the pK_a of the hydronium ion (H_3O^+) is -1.74 as given in Table 3.1.

3.6C Predicting the Strength of Bases

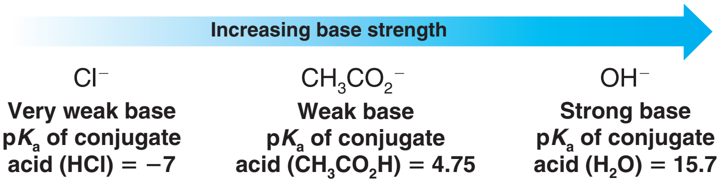
In our discussion so far we have dealt only with the strengths of acids. Arising as a natural corollary to this is a principle that allows us to estimate the **strengths of bases**. Simply stated, the principle is this:

- **The stronger the acid, the weaker will be its conjugate base.**

We can, therefore, **relate the strength of a base to the pK_a of its conjugate acid**.

- **The larger the pK_a of the conjugate acid, the stronger is the base.**

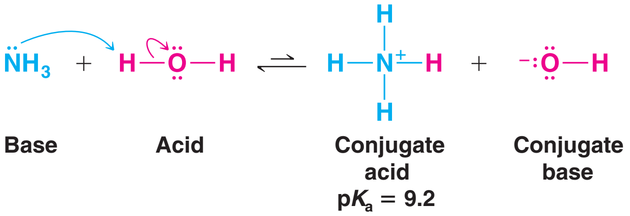
Consider the following as examples:



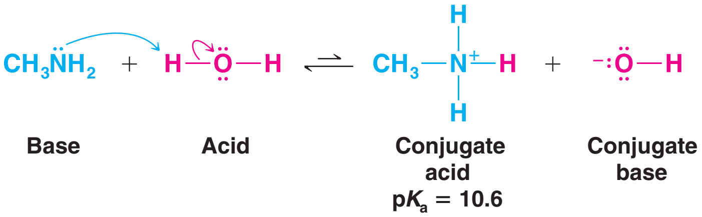
Cl^-	$CH_3CO_2^-$	OH^-
Very weak base **pK_a of conjugate acid (HCl) = −7**	**Weak base** **pK_a of conjugate acid (CH_3CO_2H) = 4.75**	**Strong base** **pK_a of conjugate acid (H_2O) = 15.7**

We see that the hydroxide ion is the strongest in this series of three bases because its conjugate acid, water, is the weakest acid. (We know that water is the weakest acid because it has the largest pK_a.)

Amines are like ammonia in that they are weak bases. Dissolving ammonia in water brings about the following equilibrium:



Dissolving methylamine in water causes the establishment of a similar equilibrium.

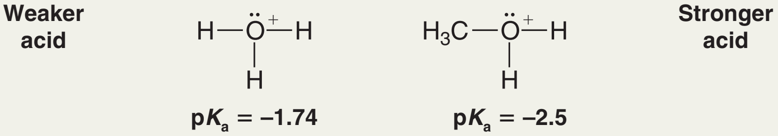


Again we can relate the basicity of these substances to the strength of their conjugate acids. The conjugate acid of ammonia is the ammonium ion, NH_4^+. The pK_a of the ammonium ion is 9.2. The conjugate acid of methylamine is the $CH_3NH_3^+$ ion. This ion, called the methylaminium ion, has pK_a = 10.6. Since the conjugate acid of methylamine is a weaker acid than the conjugate acid of ammonia, we can conclude that methylamine is a stronger base than ammonia.

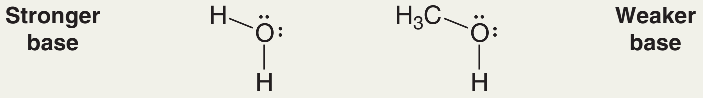
Solved Problem 3.3

Using the pK_a values in Table 3.1 decide which is the stronger base, CH_3OH or H_2O.

STRATEGY AND ANSWER From Table 3.1, we find the pK_a values of the conjugate acids of water and methanol.



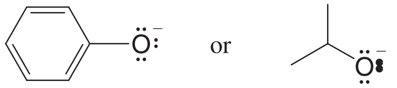
Because water is the conjugate base of the weaker acid, it is the stronger base.



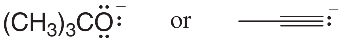
Review Problem 3.7

Using the pK_a values of analogous compounds in Table 3.1 predict which would be the stronger base.

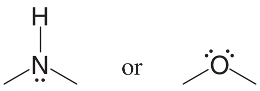
(a)



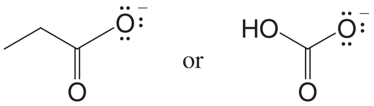
(b)



(c)



(d)



Review Problem 3.8

The pK_a of the anilinium ion ($C_6H_5\overset{+}{N}H_3$) is 4.63. On the basis of this fact, decide whether aniline ($C_6H_5NH_2$) is a stronger or weaker base than methylamine.

3.7 How to Predict the Outcome of Acid–Base Reactions

Table 3.1 gives the approximate pK_a values for a range of representative compounds. While you probably will not be expected to memorize all of the pK_a values in Table 3.1, it is a good idea to begin to learn the general order of acidity and basicity for some of the common acids and bases. The examples given in Table 3.1 are representative of their class or functional group. For example, acetic acid has a pK_a = 4.75, and carboxylic acids generally have pK_a values near this value (in the range pK_a = 3–5). Ethyl alcohol is given as an example of an alcohol, and alcohols generally have pK_a values near that of ethyl alcohol (in the pK_a range 15–18), and so on. (There are exceptions, of course, and we shall learn what these exceptions are as we go on.)

By learning the relative scale of acidity of common acids now, you will be able to predict whether or not an acid–base reaction will occur as written.

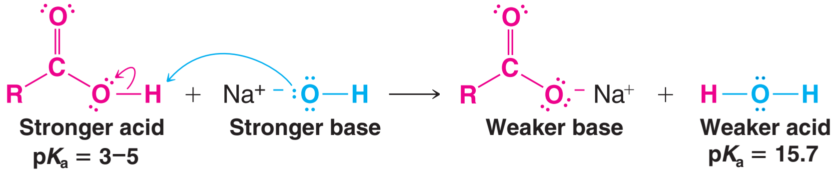
- The general principle to apply is this: **Acid–base reactions always favor the formation of the weaker acid and the weaker base**.

Helpful Hint

Formation of the weaker acid and base is an important general principle for predicting the outcome of acid–base reactions.

The reason for this is that the outcome of an acid–base reaction is determined by the position of an equilibrium. Acid–base reactions are said, therefore, to be **under equilibrium control**, and reactions under equilibrium control always favor the formation of the most stable (lowest potential energy) species. The weaker acid and weaker base are more stable (lower in potential energy) than the stronger acid and stronger base.

Using this principle, we can predict that a carboxylic acid (RCO_2H) will react with aqueous NaOH in the following way because the reaction will lead to the formation of the weaker acid (H_2O) and weaker base (RCO_2^-):



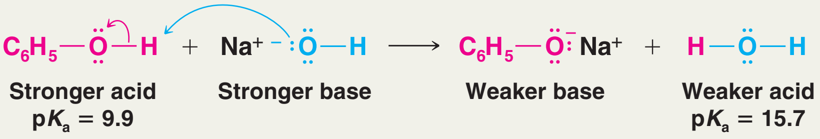
Because there is a large difference in the value of the pK_a of the two acids, the position of equilibrium will greatly favor the formation of the products. In instances like these we commonly show the reaction with a one-way arrow even though the reaction is an equilibrium.

Solved Problem 3.4

Consider the mixing of an aqueous solution of phenol, C_6H_5OH (see Table 3.1), and NaOH. What acid–base reaction, if any, would take place?

STRATEGY Consider the relative acidities of the reactant (phenol) and of the acid that might be formed (water) by a proton transfer to the base (the hydroxide ion).

ANSWER The following reaction would take place because it would lead to the formation of a weaker acid (water) from the stronger acid (phenol). It would also lead to the formation of a weaker base, C_6H_5ONa, from the stronger base, NaOH.

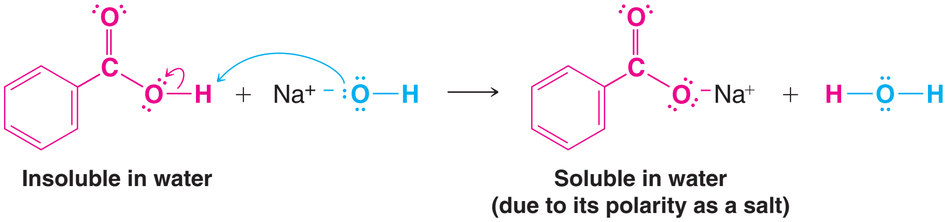


Review Problem 3.9

Predict the outcome of the following reaction.

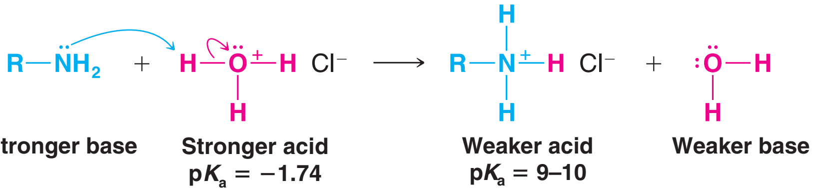
3.7A Water Solubility as the Result of Salt Formation

Although acetic acid and other carboxylic acids containing fewer than five carbon atoms are soluble in water, many other carboxylic acids of higher molecular weight are not appreciably soluble in water. Because of their acidity, however, *water-insoluble carboxylic acids dissolve in aqueous sodium hydroxide*; they do so by reacting to form water-soluble sodium salts:

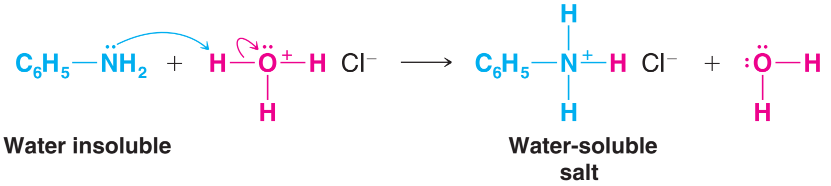


Pseudoephedrine is an amine that is sold as its hydrochloride salt.

We can also predict that an amine will react with aqueous hydrochloric acid in the following way:



While methylamine and most amines of low molecular weight are very soluble in water, amines with higher molecular weights, such as aniline ($C_6H_5NH_2$), have limited water solubility. However, these *water-insoluble amines dissolve readily in hydrochloric acid* because the acid–base reactions convert them into soluble salts:



Review Problem 3.10

Most carboxylic acids dissolve in aqueous solutions of sodium bicarbonate ($NaHCO_3$) because, as carboxylate salts, they are more polar. Write curved arrows showing the reaction between a generic carboxylic acid and sodium bicarbonate to form a carboxylate salt and H_2CO_3. (Note that H_2CO_3 is unstable and decomposes to carbon dioxide and water. You do not need to show that process.)

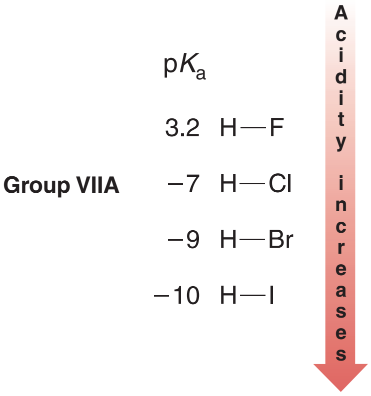
3.8 *Relationships between Structure and Acidity*

The strength of a Brønsted–Lowry acid depends on the extent to which a proton can be separated from it and transferred to a base. Removing the proton involves breaking a bond to the proton, and it involves making the conjugate base more electrically negative.

When we compare compounds in a single column of the periodic table, the strength of the bond to the proton is the dominating effect.

- Bond strength to the proton decreases as we move down the column, increasing its acidity.

This phenomenon is mainly due to decreasing effectiveness of orbital overlap between the hydrogen 1*s* orbital and the orbitals of successively larger elements in the column. The less effective the orbital overlap, the weaker is the bond, and the stronger is the acid. The acidities of the hydrogen halides furnish an example:



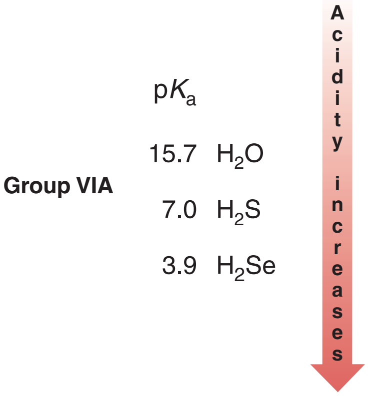
Helpful Hint

Proton acidity increases as we descend a column in the periodic table due to decreasing bond strength to the proton.

Comparing the hydrogen halides with each other, H—F is the weakest acid and H—I is the strongest. This follows from the fact that the H—F bond is by far the strongest and the H—I bond is the weakest.

Because HI, HBr, and HCl are strong acids, their conjugate bases (I^-, Br^-, Cl^-) are all weak bases. HF, however, which is less acidic than the other hydrogen halides by 10–13 orders of magnitude (compare their pK_a values), has a conjugate base that is correspondingly more basic than the other halide anions. The fluoride anion is still not nearly as basic as other species we commonly think of as bases, such as the hydroxide anion, however. A comparison of the pK_a values for HF (3.2) and H_2O (15.7) illustrates this point.

The same trend of acidities and basicities holds true in other columns of the periodic table. Consider, for example, the column headed by oxygen:



Here the strongest bond is the O—H bond and H_2O is the weakest acid; the weakest bond is the Se—H bond and H_2Se is the strongest acid.

Helpful Hint

Proton acidity increases from left to right in a given row of the periodic table due to increasing stability of the conjugate base.

- Acidity increases from left to right when we compare compounds in a given row of the periodic table.

Bond strengths vary somewhat, but the predominant factor becomes the electronegativity of the atom bonded to the hydrogen. The electronegativity of the atom in question affects acidity in two related ways. It affects the polarity of the bond to the proton and it affects the relative stability of the anion (conjugate base) that forms when the proton is lost.

We can see an example of this effect when we compare the acidities of the compounds CH_4, NH_3, H_2O, and HF. These compounds are all hydrides of first-row elements, and electronegativity increases across a row of the periodic table from left to right (see Table 1.2):

Electronegativity increases →

C N O F

Because fluorine is the most electronegative, the bond in H—F is most polarized, and the proton in H—F is the most positive. Therefore, H—F loses a proton most readily and is the most acidic in this series:

Acidity increases →

$$H_3\overset{\delta-}{C}-\overset{\delta+}{H} \qquad H_2\overset{\delta-}{N}-\overset{\delta+}{H} \qquad H\overset{\delta-}{O}-\overset{\delta+}{H} \qquad \overset{\delta-}{F}-\overset{\delta+}{H}$$

$$\mathbf{p}\boldsymbol{K}_\mathbf{a} \mathbf{= 48} \qquad \mathbf{p}\boldsymbol{K}_\mathbf{a} \mathbf{= 38} \qquad \mathbf{p}\boldsymbol{K}_\mathbf{a} \mathbf{= 15.7} \qquad \mathbf{p}\boldsymbol{K}_\mathbf{a} \mathbf{= 3.2}$$

Electrostatic potential maps for these compounds directly illustrate this trend based on electronegativity and increasing polarization of the bonds to hydrogen (Fig. 3.2). Almost no positive charge (indicated by extent of color trending toward blue) is evident at the hydrogens of methane. Very little positive charge is present at the hydrogens of ammonia. This is consistent with the weak electronegativity of both carbon and nitrogen and hence with the behavior of methane and ammonia as exceedingly weak acids (pK_a values of 48 and 38, respectively). Water shows significant positive charge at its hydrogens (pK_a more than 20 units lower than ammonia), and hydrogen fluoride clearly has the highest amount of positive charge at its hydrogen (pK_a of 3.2), resulting in strong acidity.

Because H—F is the strongest acid in this series, its conjugate base, the fluoride ion (F^-), will be the weakest base. Fluorine is the most electronegative atom and it accommodates the negative charge most readily:

← **Basicity increases**

CH_3^- H_2N^- HO^- F^-

The methanide ion (CH_3^-) is the least stable anion of the four, because carbon being the least electronegative element is least able to accept the negative charge. The methanide ion, therefore, is the strongest base in this series. [The methanide ion, a **carbanion**, and the amide ion (NH_2^-) are exceedingly strong bases because they are the conjugate bases of extremely weak acids. We shall discuss some uses of these powerful bases in Section 3.15.]

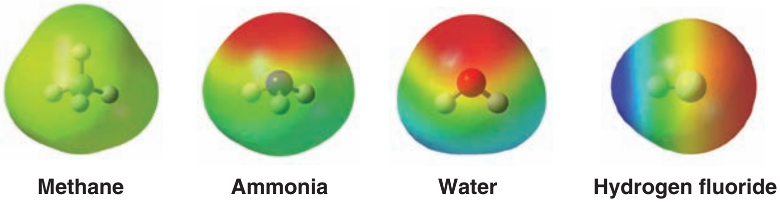


Figure 3.2 The effect of increasing electronegativity among elements from left to right in the first row of the periodic table is evident in these maps of electrostatic potential for methane, ammonia, water, and hydrogen fluoride.

Trends in acidity within the periodic table are summarized in Fig. 3.3.

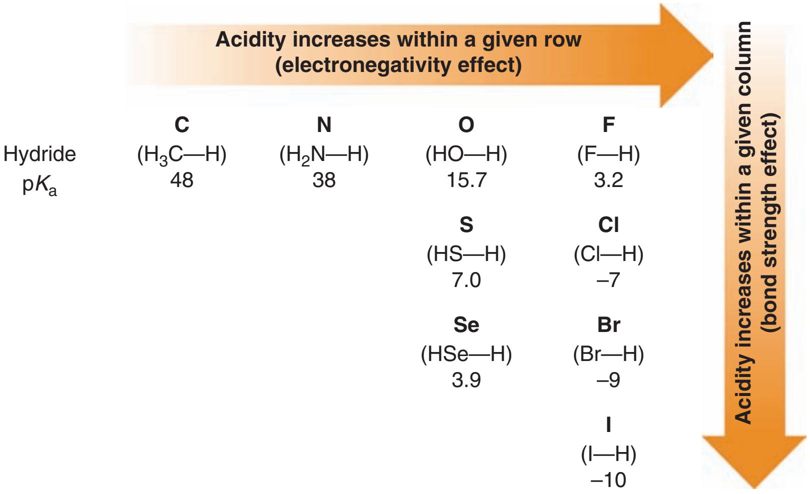
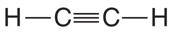


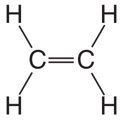
Figure 3.3 A summary of periodic trends in relative acidity. Acidity increases from left to right across a given row (electronegativity effect) and from top to bottom in a given column (bond strength effect) of the periodic table.

3.8A The Effect of Hybridization

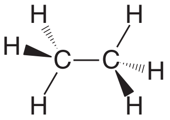
The protons of ethyne are more acidic than those of ethene, which in turn are more acidic than those of ethane:



Ethyne
pK_a = 25



Ethene
pK_a = 44



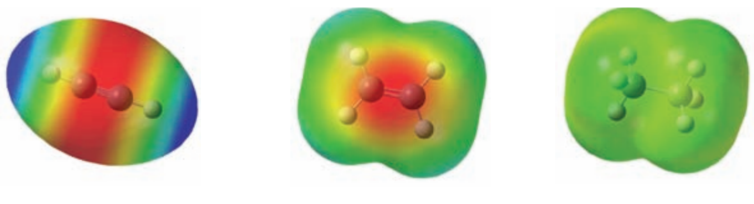
Ethane
pK_a = 50

We can explain this order of acidities on the basis of the hybridization state of carbon in each compound. Electrons of 2*s* orbitals have lower energy than those of 2*p* orbitals because *electrons in 2s orbitals tend, on the average, to be much closer to the nucleus than electrons in 2p orbitals*. (Consider the shapes of the orbitals: 2*s* orbitals are spherical and centered on the nucleus; 2*p* orbitals have lobes on either side of the nucleus and are extended into space.)

- With hybrid orbitals, **having more *s* character means that the electrons of the anion will, on the average, be lower in energy, and the anion will be more stable**.

The *sp* orbitals of the C—H bonds of ethyne have 50% *s* character (because they arise from the combination of one *s* orbital and one *p* orbital), those of the sp^2 orbitals of ethene have 33.3% *s* character, while those of the sp^3 orbitals of ethane have only 25% *s* character. This means, in effect, that the *sp* carbon atoms of ethyne act as if they were more electronegative than the sp^2 carbon atoms of ethene and the sp^3 carbon atoms of ethane. (Remember: Electronegativity measures an atom's ability to hold bonding electrons close to its nucleus, and having electrons closer to the nucleus makes it more stable.)

The effect of hybridization on acidity is borne out in the calculated electrostatic potential maps for ethyne, ethene, and ethane shown in Fig. 3.4. Some positive charge (indicated



Ethyne **Ethene** **Ethane**

Figure 3.4 Electrostatic potential maps for ethyne, ethene, and ethane.

by blue color) is clearly evident on the hydrogens of ethyne ($pK_a = 25$), but almost no positive charge is present on the hydrogens of ethene and ethane (both having pK_a values more than 20 units greater than ethyne). This is consistent with the effectively greater electronegativity of the *sp* orbitals in ethyne, which have more *s* character than the sp^2 and sp^3 orbitals in ethene and ethane. [Also evident in Fig. 3.4 is the negative charge resulting from electron density in the π bonds of ethyne and ethene (indicated by red in the region of their respective π bonds). Note the cylindrical symmetry of π electron density in the triple bond of ethyne. In the π bond of ethene there is a region of high electron density on its underneath face complementary to that visible on the top face of its double bond.]

Now we can see how the order of relative acidities of ethyne, ethene, and ethane parallels the effective electronegativity of the carbon atom in each compound:

Relative Acidity of the Hydrocarbons

$$HC{\equiv}CH > H_2C{=}CH_2 > H_3C{-}CH_3$$

Being the most electronegative, the *sp*-hybridized carbon atom of ethyne polarizes its C—H bonds to the greatest extent, causing its hydrogens to be most positive. Therefore, ethyne donates a proton to a base more readily. And, in the same way, the ethynide ion is the weakest base because the more electronegative carbon of ethyne is best able to stabilize the negative charge.

Relative Basicity of the Carbanions

$$H_3C{-}CH_2{:}^- > H_2C{=}CH{:}^- > HC{\equiv}C{:}^-$$

Notice that the explanation given here involves electronegativity, just as that given earlier to account for the relative acidities of HF, H_2O, NH_3, and CH_4.

3.8B Inductive Effects

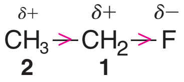
The carbon–carbon bond of ethane is completely nonpolar because at each end of the bond there are two equivalent methyl groups:

$$CH_3{-}CH_3$$

Ethane

The C—C bond is nonpolar.

This is not the case with the carbon–carbon bond of ethyl fluoride, however:



$$\overset{\delta+}{\underset{2}{CH_3}} \rightarrow \overset{\delta+}{\underset{1}{CH_2}} \rightarrow \overset{\delta-}{F}$$

One end of the bond, the one nearer the fluorine atom, is more negative than the other. This polarization of the carbon–carbon bond results from an intrinsic electron-attracting ability of the fluorine (because of its electronegativity) that is transmitted *through space* and *through the bonds of the molecule*. Chemists call this kind of effect an **inductive effect**.

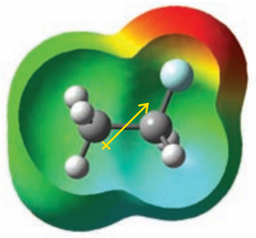


Figure 3.5 Ethyl fluoride, showing its dipole moment inside a cutaway view of the electrostatic potential at its van der Waals surface.

- **Inductive effects** are electronic effects transmitted through bonds. The inductive effect of a group can be **electron donating** or **electron withdrawing**. Inductive effects weaken as the distance from the group increases.

In the case of ethyl fluoride, the positive charge that the fluorine imparts to **C1** is greater than that imparted to **C2** because the fluorine is closer to **C1**.

Figure 3.5 shows the dipole moment for ethyl fluoride (fluoroethane). The distribution of negative charge around the electronegative fluorine is plainly evident in the calculated electrostatic potential map.

3.9 *Energy Changes*

Since we will be talking frequently about the energies of chemical systems and the relative stabilities of molecules, perhaps we should pause here for a brief review. **Energy** is defined as the capacity to do work. The two fundamental types of energy are **kinetic energy** and **potential energy**.

Kinetic energy is the energy an object has because of its motion; it equals one-half the object's mass multiplied by the square of its velocity (i.e., $\frac{1}{2}mv^2$).

Potential energy is stored energy. It exists only when an attractive or repulsive force exists between objects. Two balls attached to each other by a spring (an analogy we used for covalent bonds when we discussed infrared spectroscopy in Section 2.15) can have their potential energy increased when the spring is stretched or compressed (Fig. 3.6). If the spring is stretched, an attractive force will exist between the balls. If it is compressed, a repulsive force will exist. In either instance releasing the balls will cause the potential energy (stored energy) of the balls to be converted into kinetic energy (energy of motion).

Chemical energy is a form of potential energy. It exists because attractive and repulsive electrical forces exist between different pieces of the molecules. Nuclei attract electrons, nuclei repel each other, and electrons repel each other.

It is usually impractical (and often impossible) to describe the *absolute* amount of potential energy contained by a substance. Thus we usually think in terms of its *relative potential energy*. We say that one system has *more* or *less* potential energy than another.

Another term that chemists frequently use in this context is the term **stability** or **relative stability**. *The relative stability of a system is inversely related to its relative potential energy*.

- **The *more* potential energy an object has, the *less stable* it is.**

Consider, as an example, the relative potential energy and the relative stability of snow when it lies high on a mountainside and when it lies serenely in the valley below. Because of the attractive force of gravity, the snow high on the mountain *has greater potential energy and is much less stable* than the snow in the valley. This greater potential energy of the snow on the mountainside can become converted into the enormous kinetic energy of an avalanche. By contrast, the snow in the valley, with its lower potential energy and with its greater stability, is incapable of releasing such energy.

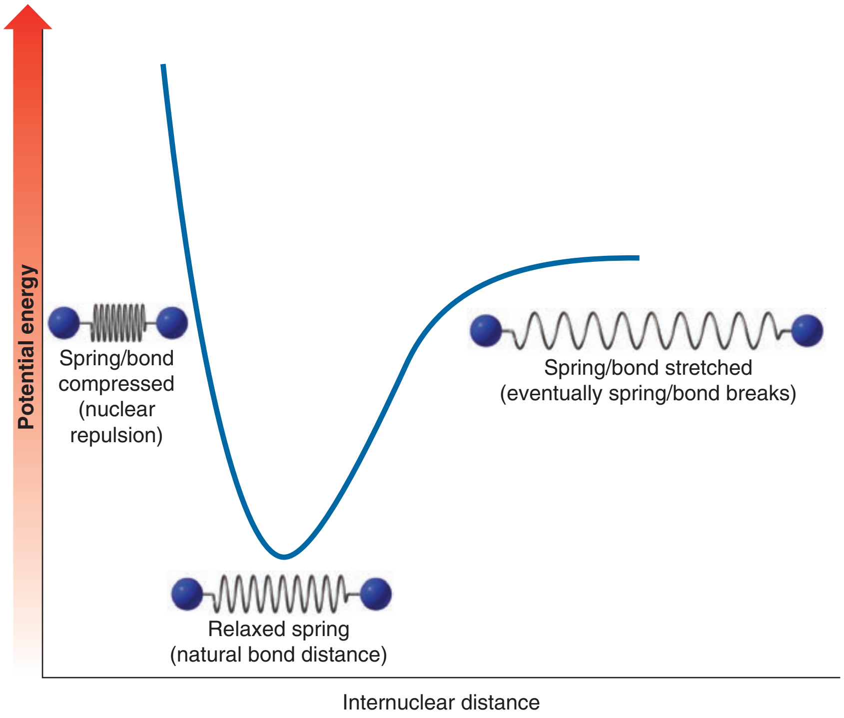


Figure 3.6 Potential energy exists between objects that either attract or repel each other. In the case of atoms joined by a covalent bond, or objects connected by a spring, the lowest potential energy state occurs when atoms are at their ideal internuclear distance (bond length), or when a spring between objects is relaxed. Lengthening or shortening the bond distance, or compressing or stretching a spring, raises the potential energy.

3.9A Potential Energy and Covalent Bonds

Atoms and molecules possess potential energy—often called chemical energy—that can be released as heat when they react. Because heat is associated with molecular motion, this release of heat results from a change from potential energy to kinetic energy.

From the standpoint of covalent bonds, the state of greatest potential energy is the state of free atoms, the state in which the atoms are not bonded to each other at all. This is true because the formation of a chemical bond is always accompanied by the lowering of the potential energy of the atoms (cf. Fig. 1.8). Consider as an example the formation of hydrogen molecules from hydrogen atoms:

$$\text{H}\cdot + \text{H}\cdot \longrightarrow \text{H—H} \qquad \Delta H^\circ = -436 \text{ kJ mol}^{-1}{*}$$

The potential energy of the atoms decreases by 436 kJ mol^{-1} as the covalent bond forms. This potential energy change is illustrated graphically in Fig. 3.7.

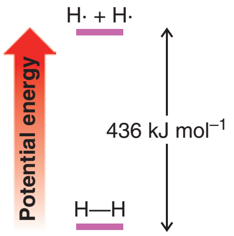


Figure 3.7 The relative potential energies of hydrogen atoms and a hydrogen molecule.

A convenient way to represent the relative potential energies of molecules is in terms of their relative **enthalpies**, or **heat contents**, *H*. (*Enthalpy* comes from *en* + *thalpein*, Greek: to heat.) The difference in relative enthalpies of reactants and products in a chemical change is called the **enthalpy change** and is symbolized by ΔH°. [The Δ (delta) in front of a quantity usually means the difference, or change, in the quantity. The superscript ° indicates that the measurement is made under standard conditions.]

By convention, the sign of ΔH° for **exothermic** reactions (those evolving heat) is negative. **Endothermic** reactions (those that absorb heat) have a positive ΔH°. The heat of reaction, ΔH°, measures the change in enthalpy of the atoms of the reactants as they are converted to products. For an exothermic reaction, the atoms have a smaller enthalpy as products than they do as reactants. For endothermic reactions, the reverse is true.

3.10 *The Relationship between the Equilibrium Constant and the Standard Free-Energy Change, ΔG°*

An important **relationship exists between the equilibrium constant (K_{eq}) and the standard free-energy change (ΔG°)** for a reaction.[†]

$$\Delta G^\circ = -RT \ln K_{eq}$$

where *R* is the gas constant and equals 8.314 J K^{-1} mol^{-1} and *T* is the absolute temperature in kelvins (K).

This equation tells us the following:

- **For a reaction to favor the formation of products when equilibrium is reached it must have a negative value for ΔG°.** Free energy must be *lost* as the reactants become products; that is, the reaction must go down an energy hill. For such a reaction the equilibrium constant will be greater than one. If ΔG° is more negative than 13 kJ mol^{-1} the equilibrium constant will be large enough for the reaction to *go to completion*, meaning that more than 99% of the reactants will be converted to products when equilibrium is reached.
- **For reactions with a positive ΔG°, the formation of products at equilibrium is unfavorable**. The equilibrium constant for these reactions will be less than one.

---

[*] The unit of energy in SI units is the joule, J, and 1 cal = 4.184 J. (Thus 1 kcal = 4.184 kJ.) A kilocalorie of energy (1000 cal) is the amount of energy in the form of heat required to raise by 1°C the temperature of 1 kg (1000 g) of water at 15°C.

[†] By standard free-energy change (ΔG°), we mean that the products and reactants are taken as being in their standard states (1 atm of pressure for a gas and 1*M* for a solution). The free-energy change is often called the **Gibbs free-energy change**, to honor the contributions to thermodynamics of J. Willard Gibbs, a professor of mathematical physics at Yale University in the latter part of the nineteenth century.

The free-energy change (ΔG°) has two components, the **enthalpy change** (ΔH°) and the **entropy change (ΔS°)**. The relationship between these three thermodynamic quantities is

$$\Delta G^\circ = \Delta H^\circ - T\,\Delta S^\circ$$

We have seen (Section 3.9) that ΔH° is associated with changes in bonding that occur in a reaction. If, collectively, stronger bonds are formed in the products than existed in the starting materials, then ΔH° will be negative (i.e., the reaction is *exothermic*). If the reverse is true, then ΔH° will be positive (the reaction is *endothermic*). **A negative value for ΔH°, therefore, will contribute to making ΔG° negative and will consequently favor the formation of products**. For the ionization of an acid, the less positive or more negative the value of ΔH°, the stronger the acid will be.

Entropy changes have to do with *changes in the relative order of a system*. **The more random a system is, the greater is its entropy**. Therefore, a positive entropy change ($+\Delta S^\circ$) is always associated with a change from a more ordered system to a less ordered one. A negative entropy change ($-\Delta S^\circ$) accompanies the reverse process. In the equation $\Delta G^\circ = \Delta H^\circ - T\,\Delta S^\circ$, the entropy change (multiplied by T) is preceded by a negative sign; this means that **a positive entropy change (from order to disorder) makes a negative contribution to ΔG° and is energetically favorable for the formation of products**.

For many reactions in which the number of molecules of products equals the number of molecules of reactants (e.g., when two molecules react to produce two molecules), the entropy change will be small. This means that except at high temperatures (where the term $T\,\Delta S^\circ$ becomes large even if ΔS° is small) the value of ΔH° will largely determine whether or not the formation of products will be favored. If ΔH° is large and negative (if the reaction is exothermic), then the reaction will favor the formation of products at equilibrium. If ΔH° is positive (if the reaction is endothermic), then the formation of products will be unfavorable.

Review Problem 3.11

State whether you would expect the entropy change, ΔS°, to be positive, negative, or approximately zero for each of the following reactions. (Assume the reactions take place in the gas phase.)

(a) $A + B \longrightarrow C$ **(b)** $A + B \longrightarrow C + D$ **(c)** $A \longrightarrow B + C$

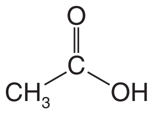
Review Problem 3.12

(a) What is the value of ΔG° for a reaction where $K_{eq} = 1$? **(b)** Where $K_{eq} = 10$? (The change in ΔG° required to produce a 10-fold increase in the equilibrium constant is a useful term to remember.) **(c)** Assuming that the entropy change for this reaction is negligible (or zero), what change in ΔH° is required to produce a 10-fold increase in the equilibrium constant?

3.11 *The Acidity of Carboxylic Acids*

Carboxylic acids are weak acids, typically having pK_a values in the range of 3–5. Alcohols, by comparison, have pK_a values in the range of 15–18, and essentially do not give up a proton unless exposed to a very strong base.

To investigate the reasons for this difference, let's consider acetic acid and ethanol as representative examples of simple carboxylic acids and alcohols.



$CH_3C(=O)OH$

Acetic acid
pK_a = 4.75
ΔG° = 27 kJ mol^{-1}

$CH_3CH_2{-}OH$

Ethanol
pK_a = 16
ΔG° = 90.8 kJ mol^{-1}

(ΔG° values are for OH proton ionization.)

Using the pK_a for acetic acid (4.75), one can calculate (Section 3.10) that the free-energy change (ΔG°) for ionization of the carboxyl proton of acetic acid is +27 kJ mol^{-1}, a moderately endergonic (unfavorable) process, since the ΔG° value is positive. Using the pK_a of ethanol (16), one can calculate that the corresponding free-energy change for ionization of

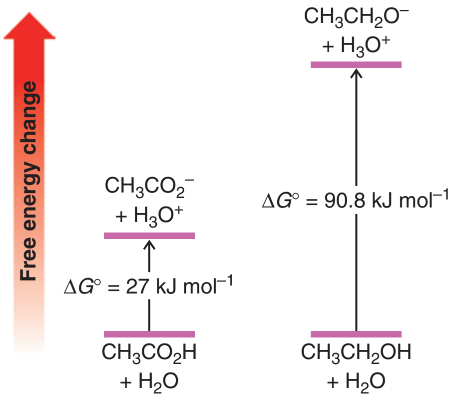
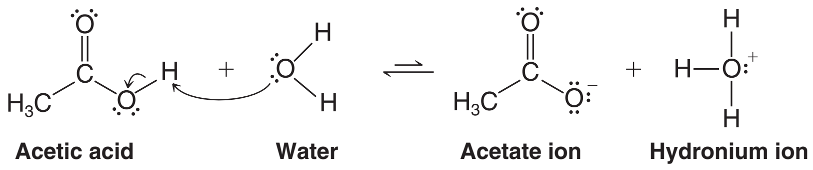


Figure 3.8 A diagram comparing the free-energy changes that accompany ionization of acetic acid and ethanol. Ethanol has a larger positive free-energy change and is a weaker acid because its ionization is more unfavorable.

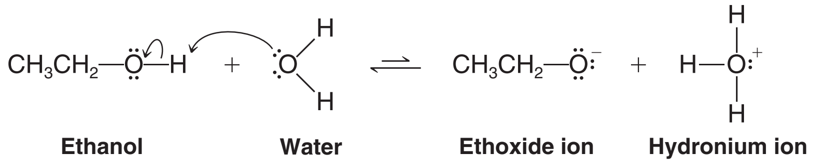
the hydroxyl proton of ethanol is $+90.8$ kJ mol^{-1}, a much more endergonic (and hence even less favorable) process. These calculations reflect the fact that ethanol is much less acidic than acetic acid. Figure 3.8 depicts the magnitude of these energy changes in a relative sense.

How do we explain the much greater acidity of carboxylic acids than alcohols? Consider first the structural changes that occur if both acetic acid and ethanol act as acids by donating a proton to water.

Acetic Acid Acting as an Acid



Ethanol Acting as an Acid

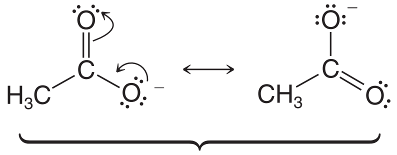


What we need to focus on is the relative stability of the conjugate bases derived from a carboxylic acid and an alcohol. This is because the smaller free-energy change for ionization of a carboxylic acid (e.g., acetic acid) as compared to an alcohol (e.g., ethanol) has been attributed to greater stabilization of the negative charge in the carboxylate ion as compared to an alkoxide ion. Greater stabilization of the carboxylate ion appears to arise from two factors: (a) delocalization of charge (as depicted by resonance structures for the carboxylate ion, Section 3.11A), and (b) an inductive electron-withdrawing effect (Section 3.8B).

3.11A The Effect of Delocalization

Delocalization of the negative charge is possible in a carboxylate anion, but it is not possible in an alkoxide ion. We can show how delocalization is possible in carboxylate ions by writing resonance structures for the acetate ion.

Two resonance structures that can be written for acetate anion



Resonance stabilization in acetate ion
(The structures are equivalent and there is no requirement for charge separation.)

The two resonance structures we drew above distributed the negative charge to both oxygen atoms of the carboxylate group, thereby stabilizing the charge. This is a **delocalization effect** (by resonance). In contrast, no resonance structures are possible for an alkoxide ion, such as ethoxide. (You may wish to review the rules we have given in Section 1.8 for writing proper resonance structures.)

$$CH_3-CH_2-\ddot{\underset{\cdot\cdot}{O}}-H \quad + \quad H_2O \quad \rightleftharpoons \quad CH_3-CH_2-\ddot{\underset{\cdot\cdot}{O}}\!:^- \quad + \quad H_3O^+$$

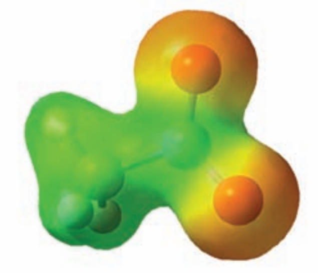
No resonance stabilization **No resonance stabilization**

No resonance structures can be drawn for either ethanol or ethoxide anion.

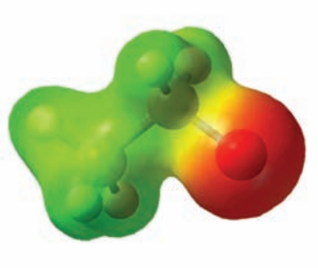
A rule to keep in mind is that **charge delocalization is always a stabilizing factor**, and because of charge stabilization, the energy difference for formation of a carboxylate ion from a carboxylic acid is less than the energy difference for formation of an alkoxide ion from an alcohol. Since the energy difference for ionization of a carboxylic acid is less than for an alcohol, the carboxylic acid is a stronger acid.

3.11B The Inductive Effect

We have already shown how the negative charge in a carboxylate ion can be delocalized over two oxygen atoms by resonance. However, the electronegativity of these oxygen atoms further helps to stabilize the charge, by what is called an **inductive electron-withdrawing effect**. A carboxylate ion has two oxygen atoms whose combined electronegativity stabilizes the charge more than in an alkoxide ion, which has only a single electronegative oxygen atom. In turn, this lowers the energy barrier to forming the carboxylate ion, making a carboxylic acid a stronger acid than an alcohol. This effect is evident in electrostatic potential maps depicting approximating the bonding electron density for the two anions (Fig. 3.9). Negative charge in the acetate anion is evenly distributed over the two oxygen atoms, whereas in ethoxide the negative charge is localized on its sole oxygen atom (as indicated by red in the electrostatic potential map).



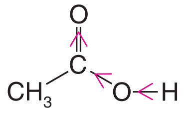
Acetate anion



Ethoxide anion

Figure 3.9 Calculated electrostatic potential maps at a surface approximating the bonding electron density for acetate anion and ethoxide anion. Although both molecules carry the same −1 net charge, acetate stabilizes the charge better by dispersing it over both oxygen atoms.

It is also reasonable to expect that a carboxylic acid would be a stronger acid than an alcohol when considering each as a neutral molecule (i.e., prior to loss of a proton), because both functional groups have a highly polarized O—H bond, which in turn weakens the bond to the hydrogen atom. However, the significant electron-withdrawing effect of the carbonyl group in acetic acid and the absence of an adjacent electron-withdrawing group in ethanol make the carboxylic acid hydrogen much more acidic than the alcohol hydrogen.



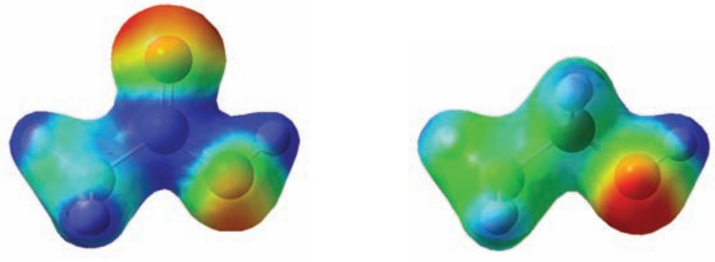
$$CH_3-\overset{\overset{O}{\Uparrow}}{C}\leftarrow O\leftarrow H$$

Acetic acid (stronger acid)

$$CH_3-CH_2-O\leftarrow H$$

Ethanol (weaker acid)

Electrostatic potential maps at approximately the bond density surface for acetic acid and ethanol (Fig. 3.10) clearly show the positive charge at the carbonyl carbon of acetic acid, as compared to the CH_2 carbon of ethanol.



Acetic acid **Ethanol**

Figure 3.10 Maps of electrostatic potential at approximately the bond density surface for acetic acid and ethanol. The positive charge at the carbonyl carbon of acetic acid is evident in the blue color of the electrostatic potential map at that position, as compared to the hydroxyl carbon of ethanol. The inductive electron-withdrawing effect of the carbonyl group in carboxylic acids contributes to the acidity of this functional group.

3.11C Summary and a Comparison of Conjugate Acid–Base Strengths

In summary, the greater acidity of a carboxylic acid is predominantly due to the ability of its conjugate base (a carboxylate ion) to stabilize a negative charge better than an alkoxide ion, the conjugate base of an alcohol. In other words, the conjugate base of a carboxylic acid is a weaker base than the conjugate base of an alcohol. Therefore, since there is an inverse strength relationship between an acid and its conjugate base, a carboxylic acid is a stronger acid than an alcohol.

Helpful Hint

The more stable a conjugate base is, the stronger the corresponding acid.

Review Problem 3.13

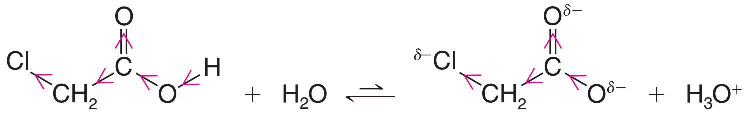
Draw contributing resonance structures and a hybrid resonance structure that explain two related facts: The carbon–oxygen bond distances in the acetate ion are the same, and the oxygens of the acetate ion bear equal negative charges.

3.11D Inductive Effects of Other Groups

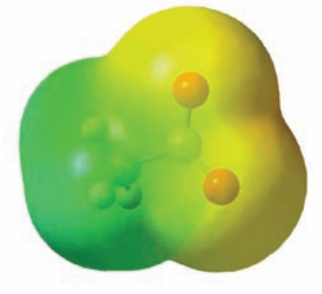
The acid-strengthening effect of other electron-attracting groups (other than the carbonyl group) can be shown by comparing the acidities of acetic acid and chloroacetic acid:

pK_a = 4.75 **pK_a = 2.86**

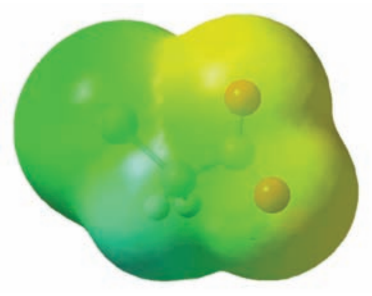
This is an example of a **substituent effect**. The greater acidity of chloroacetic acid can be attributed, in part, to the extra electron-attracting inductive effect of the electronegative chlorine atom. By adding its inductive effect to that of the carbonyl group and the oxygen, it makes the hydroxyl proton of chloroacetic acid even more positive than that of acetic acid. It also stabilizes the chloroacetate ion that is formed when the proton is lost *by dispersing its negative charge* (Fig. 3.11):



Dispersal of charge always makes a species more stable, and, as we have seen now in several instances, **any factor that stabilizes the conjugate base of an acid increases the strength of the acid**. (In Section 3.12, we shall see that entropy changes in the solvent are also important in explaining the increased acidity of chloroacetic acid.)



Acetate anion



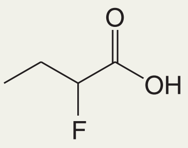
Chloroacetate anion

Figure 3.11 The electrostatic potential maps for acetate and chloroacetate ions show the relatively greater ability of chloroacetate to disperse the negative charge.

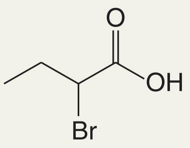
Solved Problem 3.5

Which compound in each pair would be most acidic?

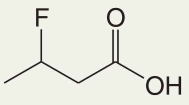
(a)



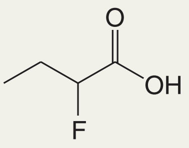
or



(b)



or



STRATEGY AND ANSWER Decide what is similar in each pair and what is different. In pair (a), the difference is the halogen substituent on carbon 2. In the first example it is fluorine; in the second it is bromine. Fluorine is much more electronegative (electron-attracting) than bromine (Table 1.2); therefore it will be able to disperse the negative charge of the anion formed when the proton is lost. Thus the first compound will be the stronger acid. In pair (b), the difference is the position of the fluorine substituents. In the second compound the fluorine is closer to the carboxyl group where it will be better able to disperse the negative charge in the anion formed when the proton is lost. The second compound will be the stronger acid.

Which would you expect to be the stronger acid? Explain your reasoning in each instance.

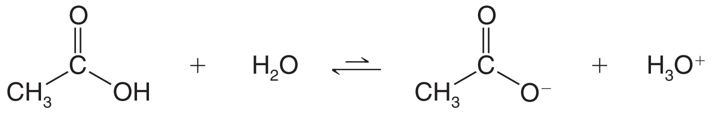
Review Problem 3.14

(a) CH_2ClCO_2H or $CHCl_2CO_2H$ **(c)** CH_2FCO_2H or CH_2BrCO_2H

(b) CCl_3CO_2H or $CHCl_2CO_2H$ **(d)** CH_2FCO_2H or $CH_2FCH_2CO_2H$

3.12 *The Effect of the Solvent on Acidity*

In the absence of a solvent (i.e., in the gas phase), most acids are far weaker than they are in solution. In the gas phase, for example, acetic acid is estimated to have a pK_a of about 130 (a K_a of $\sim 10^{-130}$)! The reason is this: When an acetic acid molecule donates a proton to a water molecule in the gas phase, the ions that are formed are oppositely charged particles and the particles must become separated:



In the absence of a solvent, separation is difficult. In solution, solvent molecules surround the ions, insulating them from one another, stabilizing them, and making it far easier to separate them than in the gas phase.

In a solvent such as water, called a protic solvent, solvation by hydrogen bonding is important (Section 2.13D).

- A **protic solvent** is one that has a hydrogen atom attached to a strongly electronegative element such as oxygen or nitrogen.

Molecules of a protic solvent, therefore, can form hydrogen bonds to the unshared electron pairs of oxygen (or nitrogen) atoms of an acid and its conjugate base, but they may not stabilize both equally.

Consider, for example, the ionization of acetic acid in aqueous solution. Water molecules solvate both the undissociated acid (CH_3CO_2H) and its anion ($CH_3CO_2^-$) by forming hydrogen bonds to them (as shown for hydroxide in Section 3.2B). However, hydrogen bonding to $CH_3CO_2^-$ is much stronger than to CH_3CO_2H because the water molecules are more attracted by the negative charge. This differential solvation, moreover, has important consequences for the entropy change that accompanies the ionization. **Solvation of any species decreases the entropy of the solvent because the solvent molecules become much more ordered as they surround molecules of the solute**. Because solvation of $CH_3CO_2^-$ is stronger, the solvent molecules become more orderly around it. The entropy change ($\Delta S°$) for the ionization of acetic acid, therefore, is negative. This means that the $T\,\Delta S°$ term in the equation $\Delta G° = \Delta H° - T\,\Delta S°$ makes an acid-weakening positive contribution to $\Delta G°$. In fact, as Table 3.2 shows, the $T\,\Delta S°$ term contributes more to the value of $\Delta G°$ than does $\Delta H°$, and accounts for the fact that the free-energy change for the ionization of acetic acid is positive (unfavorable).

We saw in Section 3.11D that chloroacetic acid is a stronger acid than acetic acid, and we attributed this increased acidity to the presence of the electron-withdrawing chlorine atom. Table 3.2 shows us that both $\Delta H°$ and $T\,\Delta S°$ are more favorable for the ionization of chloroacetic acid ($\Delta H°$ is more negative by 4.2 kJ mol^{-1}, and $T\,\Delta S°$ is less negative by 7 kJ mol^{-1}). The larger contribution is clearly in the entropy term. Apparently, by stabilizing the chloroacetate anion, the chlorine atom makes the chloroacetate ion less prone to cause an ordering of the solvent because it requires less stabilization through solvation.

TABLE 3.2 Thermodynamic Values for the Dissociation of Acetic and Chloroacetic Acids in H_2O at 25°C

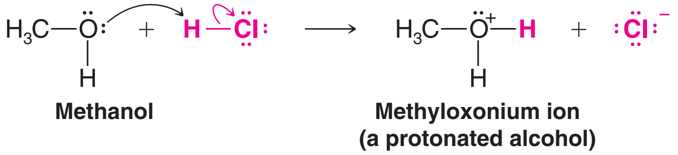
Acid	pK_a	$\Delta G°$ (kJ mol^{-1})	=	$\Delta H°$ (kJ mol^{-1})	−	$T\,\Delta S°$ (kJ mol^{-1})
CH_3CO_2H	4.75	+27		−0.4		−28
$ClCH_2CO_2H$	2.86	+16		−4.6		−21

Reprinted with permission of John Wiley & Sons, Inc. from March, J., *Advanced Organic Chemistry, Fourth Edition*, p. 272. Copyright 1992.

3.13 *Organic Compounds as Bases*

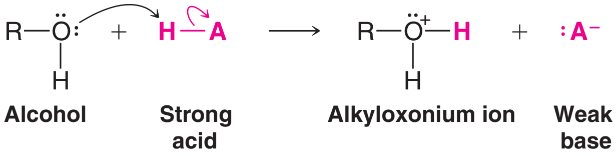
If an organic compound contains an atom with an unshared electron pair, it is a potential base. We saw in Section 3.6C that compounds with an unshared electron pair on a nitrogen atom (i.e., amines) act as bases. Let us now consider several examples in which organic compounds having an unshared electron pair on an oxygen atom act in the same way.

Dissolving gaseous HCl in methanol brings about an acid–base reaction much like the one that occurs with water (Section 3.2A):

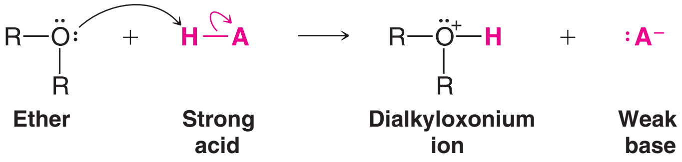


The conjugate acid of the alcohol is often called a **protonated alcohol**, although more formally it is called an **alkyloxonium ion** or simply an **oxonium ion**.

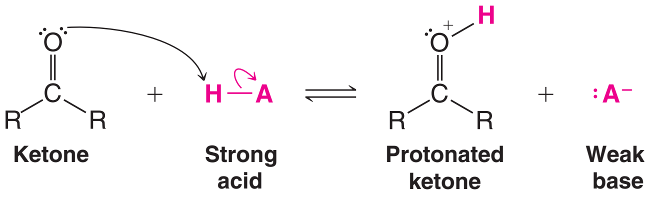
Alcohols, in general, undergo this same reaction when they are treated with solutions of strong acids such as HCl, HBr, HI, and H_2SO_4:



So, too, do ethers:



Compounds containing a carbonyl group also act as bases in the presence of a strong acid:

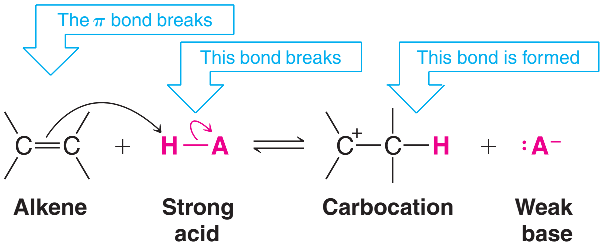


Helpful Hint

Proton transfers are a common first step in many reactions we shall study.

Proton transfer reactions like these are often the first step in many reactions that alcohols, ethers, aldehydes, ketones, esters, amides, and carboxylic acids undergo. The pK_a values for some of these protonated intermediates are given in Table 3.1.

An atom with an unshared electron pair is not the only locus that confers basicity on an organic compound. The π bond of an alkene can have the same effect. Later we shall study many reactions in which, as a first step, alkenes react with a strong acid by accepting a proton in the following way:



In this reaction the electron pair of the π bond of the alkene is used to form a bond between one carbon of the alkene and the proton donated by the strong acid. Notice that two bonds are broken in this process: the π bond of the double bond and the bond between the proton of the acid and its conjugate base. One new bond is formed: a bond between a carbon of the alkene and the proton. This process leaves the other carbon of the alkene trivalent, electron deficient, and with a formal positive charge. A compound containing a carbon of this type is called a **carbocation** (Section 3.4). As we shall see in later chapters, carbocations are unstable intermediates that react further to produce stable molecules.

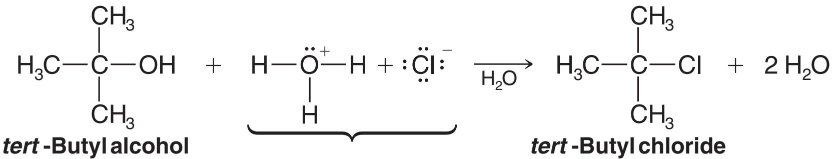
Review Problem 3.15

It is a general rule that any organic compound containing oxygen, nitrogen, or a multiple bond will dissolve in concentrated sulfuric acid. Explain the basis of this rule in terms of acid–base reactions and intermolecular forces.

3.14 *A Mechanism for an Organic Reaction*

In Chapter 6 we shall begin our study of organic **reaction mechanisms** in earnest. Let us consider now one mechanism as an example, one that allows us to apply some of the chemistry we have learned in this chapter and one that, at the same time, will reinforce what we have learned about how curved arrows are used to illustrate mechanisms.

Dissolving *tert*-butyl alcohol in concentrated (concd) aqueous hydrochloric acid soon results in the formation of *tert*-butyl chloride. The reaction is a substitution reaction:



$$(CH_3)_3C{-}OH + \underbrace{H_3O^+ + Cl^-}_{\text{Concd HCl}} \xrightarrow[H_2O]{} (CH_3)_3C{-}Cl + 2\,H_2O$$

***tert*-Butyl alcohol (soluble in H_2O)**

Concd HCl

***tert*-Butyl chloride (insoluble in H_2O)**

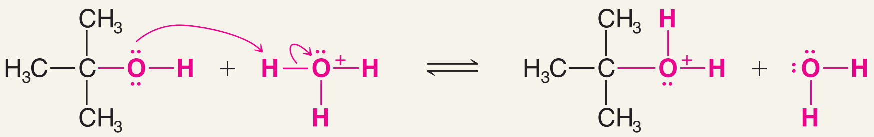
That a reaction has taken place is obvious when one actually does the experiment. *tert*-Butyl alcohol is soluble in the aqueous medium; however, *tert*-butyl chloride is not, and consequently it separates from the aqueous phase as another layer in the flask. It is easy to remove this nonaqueous layer, purify it by distillation, and thus obtain the *tert*-butyl chloride.

Considerable evidence, described later, indicates that the reaction occurs in the following way.

A MECHANISM FOR THE REACTION

Reaction of *tert*-Butyl Alcohol with Concentrated Aqueous HCl

Step 1



$$(CH_3)_3C{-}\ddot{O}{-}H + H{-}\overset{+}{O}H_2 \rightleftharpoons (CH_3)_3C{-}\overset{+}{O}H_2 + H{-}\ddot{O}{-}H$$

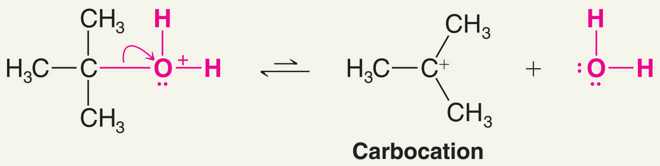
***tert*-Butyloxonium ion**

***tert*-Butyl alcohol acts as a base and accepts a proton from the hydronium ion. (Chloride anions are spectators in this step of the reaction.)**

The products are a protonated alcohol and water (the conjugate acid and base).

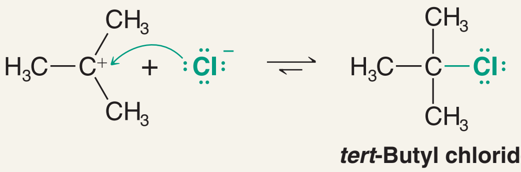
(continues on the next page)

Step 2



The bond between the carbon and oxygen of the *tert*-butyloxonium ion breaks heterolytically, leading to the formation of a carbocation and a molecule of water.

Step 3



The carbocation, acting as a Lewis acid, accepts an electron pair from a chloride ion to become the product.

Notice that **all of these steps involve acid–base reactions**. Step 1 is a straightforward Brønsted acid–base reaction in which the alcohol oxygen removes a proton from the hydronium ion. Step 2 is the reverse of a Lewis acid–base reaction. In it, the carbon–oxygen bond of the protonated alcohol breaks heterolytically as a water molecule departs with the electrons of the bond. This happens, in part, because the alcohol is protonated. The presence of a formal positive charge on the oxygen of the protonated alcohol weakens the carbon–oxygen bond by drawing the electrons in the direction of the positive oxygen. Step 3 is a Lewis acid–base reaction, in which a chloride anion (a Lewis base) reacts with the carbocation (a Lewis acid) to form the product.

A question might arise: Why doesn't a molecule of water (also a Lewis base) instead of a chloride ion react with the carbocation? After all, there are many water molecules around, since water is the solvent. The answer is that this step does occur sometimes, but it is simply the reverse of step 2. That is to say, not all of the carbocations that form go on directly to become product. Some react with water to become protonated alcohols again. However, these will dissociate again to become carbocations (even if, before they do, they lose a proton to become the alcohol again). Eventually, however, most of them are converted to the product because, under the conditions of the reaction, the equilibrium of the last step lies far to the right, and the product separates from the reaction mixture as a second phase.

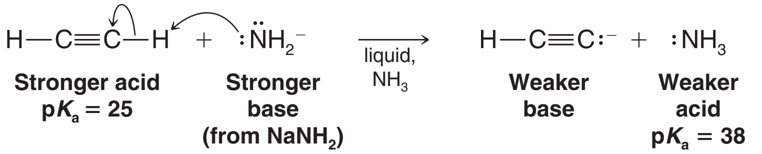
3.15 *Acids and Bases in Nonaqueous Solutions*

If you were to add sodium amide ($NaNH_2$) to water in an attempt to carry out a reaction using the amide ion (NH_2^-) as a very powerful base, the following reaction would take place immediately:

The amide ion would react with water to produce a solution containing hydroxide ion (a much weaker base) and ammonia. This example illustrates what is called the **leveling effect** of the solvent. *Water*, the solvent here, *donates a proton to any base stronger than a hydroxide ion*. Therefore, *it is not possible to use a base stronger than hydroxide ion in aqueous solution*.

We can use bases stronger than hydroxide ion, however, if we choose solvents that are weaker acids than water. We can use amide ion (e.g., from $NaNH_2$) in a solvent such as hexane, diethyl ether, or liquid NH_3 (the liquified gas, bp −33°C, not the aqueous solution that you may have used in your general chemistry laboratory). All of these solvents are very weak acids (we generally don't think of them as acids), and therefore they will not donate a proton even to the strong base NH_2^-.

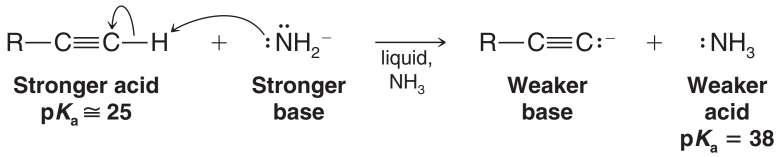
We can, for example, convert ethyne to its conjugate base, a carbanion, by treating it with sodium amide in liquid ammonia:



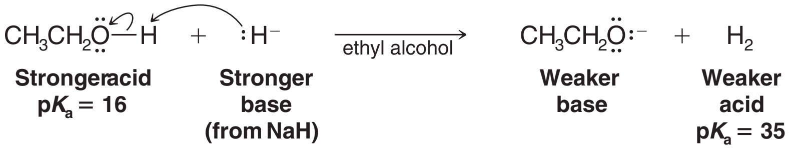
Helpful Hint

We shall use this reaction as part of our introduction to organic synthesis in Chapter 7.

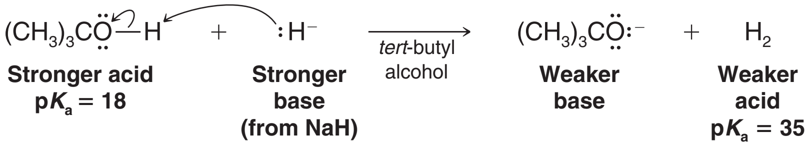
Most **terminal alkynes** (alkynes with a proton attached to a triply bonded carbon) have pK_a values of about 25; therefore, all react with sodium amide in liquid ammonia in the same way that ethyne does. The general reaction is



Alcohols are often used as solvents for organic reactions because, being somewhat less polar than water, they dissolve less polar organic compounds. Using alcohols as solvents also offers the advantage of using RO^- ions (called **alkoxide ions**) as bases. Alkoxide ions are somewhat stronger bases than hydroxide ions because alcohols are weaker acids than water. For example, we can create a solution of sodium ethoxide (CH_3CH_2ONa) in ethyl alcohol by adding sodium hydride (NaH) to ethyl alcohol. We use a large excess of ethyl alcohol because we want it to be the solvent. Being a very strong base, the hydride ion reacts readily with ethyl alcohol:



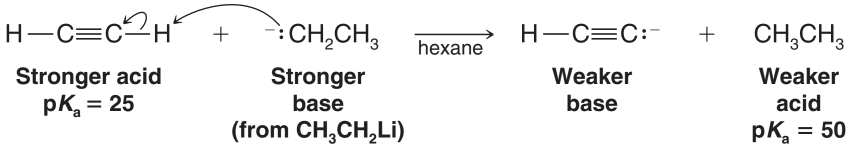
The *tert*-butoxide ion, $(CH_3)_3CO^-$, in *tert*-butyl alcohol, $(CH_3)_3COH$, is a stronger base than the ethoxide ion in ethyl alcohol, and it can be prepared in a similar way:



Although the carbon–lithium bond of an alkyllithium (RLi) has covalent character, it is polarized so as to make the carbon negative:

$$\overset{\delta-}{R}\leftarrow\overset{\delta+}{Li}$$

Alkyllithium reagents react as though they contain alkanide ($R:^-$) ions and, being the conjugate bases of alkanes, alkanide ions are the strongest bases that we shall encounter. Ethyllithium (CH_3CH_2Li), for example, acts as though it contains an ethanide ($CH_3CH_2:^-$) carbanion. It reacts with ethyne in the following way:



Alkyllithiums can be easily prepared by allowing an alkyl bromide to react with lithium metal in an ether solvent (such as diethyl ether). See Section 12.6.

Review Problem 3.16

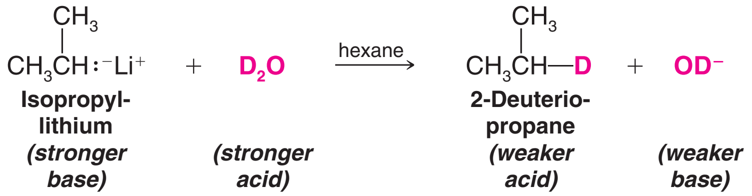
Write equations for the acid–base reaction that would occur when each of the following compounds or solutions are mixed. In each case label the stronger acid and stronger base, and the weaker acid and weaker base, by using the appropriate pK_a values (Table 3.1). (If no appreciable acid–base reaction would occur, you should indicate this.)

(a) NaH is added to CH_3OH.
(b) $NaNH_2$ is added to CH_3CH_2OH.
(c) Gaseous NH_3 is added to ethyllithium in hexane.
(d) NH_4Cl is added to sodium amide in liquid ammonia.
(e) $(CH_3)_3CONa$ is added to H_2O.
(f) $NaOH$ is added to $(CH_3)_3COH$.

3.16 *Acid–Base Reactions and the Synthesis of Deuterium- and Tritium-Labeled Compounds*

Chemists often use compounds in which deuterium or tritium atoms have replaced one or more hydrogen atoms of the compound as a method of “labeling” or identifying particular hydrogen atoms. Deuterium (^2H) and tritium (^3H) are isotopes of hydrogen with masses of 2 and 3 atomic mass units (amu), respectively.

One way to introduce a deuterium or tritium atom into a specific location in a molecule is through the acid–base reaction that takes place when a very strong base is treated with D_2O or T_2O (water that has deuterium or tritium in place of its hydrogens). For example, treating a solution containing $(CH_3)_2CHLi$ (isopropyllithium) with D_2O results in the formation of propane labeled with deuterium at the central atom:



$$\underset{\substack{\textbf{Isopropyl-}\\\textbf{lithium}\\\textit{(stronger}\\\textit{base)}}}{CH_3\overset{\displaystyle CH_3}{\overset{|}{C}}H{:}^-Li^+} + \underset{\substack{\textit{(stronger}\\\textit{acid)}}}{D_2O} \xrightarrow{\text{hexane}} \underset{\substack{\textbf{2-Deuterio-}\\\textbf{propane}\\\textit{(weaker}\\\textit{acid)}}}{CH_3\overset{\displaystyle CH_3}{\overset{|}{C}}H{-}D} + \underset{\substack{\textit{(weaker}\\\textit{base)}}}{OD^-}$$

Solved Problem 3.6

Assuming you have available propyne, a solution of sodium amide in liquid ammonia, and T_2O, show how you would prepare the tritium-labeled compound $CH_3C{\equiv}CT$.

ANSWER First add propyne to sodium amide in liquid ammonia. The following acid–base reaction will take place:

$$\underset{\substack{\textbf{Stronger}\\\textbf{acid}}}{CH_3C{\equiv}CH} + \underset{\substack{\textbf{Stronger}\\\textbf{base}}}{NH_2^-} \xrightarrow[\text{liq. ammonia}]{} \underset{\substack{\textbf{Weaker}\\\textbf{base}}}{CH_3C{\equiv}C{:}^-} + \underset{\substack{\textbf{Weaker}\\\textbf{acid}}}{NH_3}$$

Then adding T_2O (a much stronger acid than NH_3) to the solution will produce $CH_3C \equiv CT$:

$$\underset{\substack{\textbf{Stronger}\\\textbf{base}}}{CH_3C{\equiv}C{:}^-} + \underset{\substack{\textbf{Stronger}\\\textbf{acid}}}{T_2O} \xrightarrow[\text{liq. ammonia}]{} \underset{\substack{\textbf{Weaker}\\\textbf{acid}}}{CH_3C{\equiv}CT} + \underset{\substack{\textbf{Weaker}\\\textbf{base}}}{OT^-}$$

Review Problem 3.17

Complete the following acid–base reactions:

(a) $HC{\equiv}CH + NaH \xrightarrow[\text{hexane}]{}$
(b) The solution obtained in (a) + $D_2O \longrightarrow$
(c) $CH_3CH_2Li + D_2O \xrightarrow[\text{hexane}]{}$
(d) $CH_3CH_2OH + NaH \xrightarrow[\text{hexane}]{}$
(e) The solution obtained in (d) + $T_2O \longrightarrow$
(f) $CH_3CH_2CH_2Li + D_2O \xrightarrow[\text{hexane}]{}$

3.17 *Applications of Basic Principles*

Again we review how certain basic principles apply to topics we have studied in this chapter.

Electronegativity Differences Polarize Bonds In Section 3.1A we learned that *heterolysis* of a covalent bond is aided when the bond is polarized by a difference in electronegativity of the bonded atoms. We saw how this principle applies to the heterolysis of bonds to carbon in Section 3.4 and in explaining the strength of acids in Sections 3.8 and 3.11B.

Polarized Bonds Underlie Inductive Effects In Section 3.11B we saw how polarized bonds explain effects that we call *inductive effects* and how these effects are part of the explanation for why carboxylic acids are more acidic than corresponding alcohols.

Opposite Charges Attract This principle is fundamental to understanding *Lewis acid–base theory* as we saw in Section 3.3A. Positively charged centers in molecules that are electron pair acceptors are attracted to negatively charged centers in electron pair donors. In Section 3.4 we saw this principle again in the reaction of carbocations (positively charged Lewis acids) with anions (which are negatively charged by definition) and other Lewis bases.

Nature Prefers States of Lower Potential Energy In Section 3.9A we saw how this principle explains the energy changes—called *enthalpy changes*—that take place when covalent bonds form, and in Section 3.10 we saw the role enthalpy changes play in explaining how large or how small the equilibrium constant for a reaction is. The lower the potential energy of the products, the larger is the equilibrium constant, and the more favored is the formation of the products when equilibrium is reached. This section also introduced a related principle: **Nature prefers disorder to order**—or, to put it another way, *a positive entropy change* for a reaction favors the formation of the products at equilibrium.

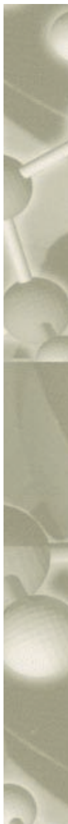
Resonance Effects Can Stabilize Molecules and Ions When a molecule or ion can be represented by two or more equivalent resonance structures, then the molecule or ion will be stabilized (will have its potential energy lowered) by delocalization of charge. In Section 3.11A we saw how this effect helps explain the greater acidity of carboxylic acids when compared to corresponding alcohols.

In This Chapter

In Chapter 3 you studied acid–base chemistry, one of the most important topics needed to learn organic chemistry. If you master acid–base chemistry you will be able to understand most of the reactions that you study in organic chemistry, and by understanding how reactions work, you will be able to learn and remember them more easily.

You have reviewed the Brønsted–Lowry definition of acids and bases and the meanings of pH and pK_a. You have learned to identify the most acidic hydrogen atoms in a molecule based on a comparison of pK_a values. You will see in many cases that Brønsted–Lowry acid–base reactions either initiate or complete an organic reaction, or prepare an organic molecule for further reaction. The Lewis definition of acids and bases may have been new to you. However, you will see over and over again that Lewis acid–base reactions which involve either the donation of an electron pair to form a new covalent bond or the departure of an electron pair to break a covalent bond are central steps in many organic reactions. The vast majority of organic reactions you will study are either Brønsted–Lowry or Lewis acid–base reactions.

Your knowledge of organic structure and polarity from Chapters 1 and 2 has been crucial to your understanding of acid–base reactions. You have seen that stabilization of charge by delocalization is key to determining how readily an acid will give up a proton, or how readily a base will accept a proton. In addition, you have learned the essential skill of drawing curved arrows to accurately show the movement of electrons in these processes. With these concepts and skills you will be prepared to understand how organic reactions occur on a step-by-step basis—something organic chemists call “a mechanism for the reaction.”



So, continue to work hard to master acid–base chemistry and other fundamentals. Your toolbox is quickly filling with the tools you need for overall success in organic chemistry!

Key Terms and Concepts

The key terms and concepts that are highlighted in **bold, blue text** within the chapter are defined in the glossary (at the back of the book) and have hyperlinked definitions in the accompanying ***WileyPLUS*** course (www.wileyplus.com).

Problems

Note to Instructors: Many of the homework problems are available for assignment via *WileyPLUS*, an online teaching and learning solution.

BRØNSTED–LOWRY ACIDS AND BASES

3.18 What is the conjugate base of each of the following acids?

(a) NH_3 **(b)** H_2O **(c)** H_2 **(d)** $HC{\equiv}CH$ **(e)** CH_3OH **(f)** H_3O^+

3.19 List the bases you gave as answers to Exercise 3.18 in order of decreasing basicity.

3.20 What is the conjugate acid of each of the following bases?

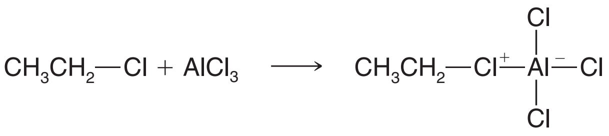
(a) HSO_4^- **(b)** H_2O **(c)** CH_3NH_2 **(d)** NH_2^- **(e)** $CH_3CH_2^-$ **(f)** $CH_3CO_2^-$

3.21 List the acids you gave as answers to Exercise 3.20 in order of decreasing acidity.

LEWIS ACIDS AND BASES

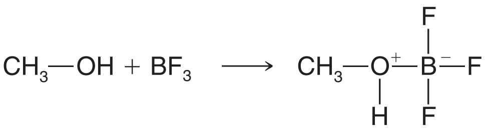
3.22 Designate the Lewis acid and Lewis base in each of the following reactions:

(a)



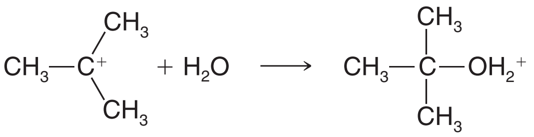
$$CH_3CH_2-Cl + AlCl_3 \longrightarrow CH_3CH_2-\overset{+}{Cl}-\overset{\displaystyle Cl}{\underset{\displaystyle Cl}{\overset{-}{Al}}}-Cl$$

(b)



$$CH_3-OH + BF_3 \longrightarrow CH_3-\underset{\displaystyle H}{\overset{+}{O}}-\overset{\displaystyle F}{\underset{\displaystyle F}{\overset{-}{B}}}-F$$

(c)



$$CH_3-\overset{\displaystyle CH_3}{\underset{\displaystyle CH_3}{C^+}} + H_2O \longrightarrow CH_3-\overset{\displaystyle CH_3}{\underset{\displaystyle CH_3}{C}}-OH_2^+$$

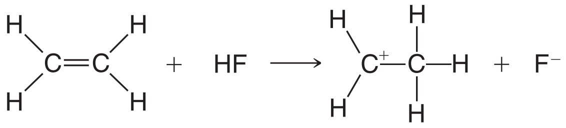
CURVED-ARROW NOTATION

3.23 Rewrite each of the following reactions using curved arrows and show all nonbonding electron pairs:

(a) $CH_3OH + HI \rightarrow CH_3OH_2^+ + I^-$

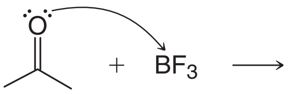
(b) $CH_3NH_2 + HCl \rightarrow CH_3NH_3^+ + Cl^-$

(c)

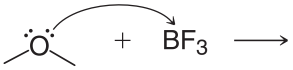


3.24 Follow the curved arrows and write the products.

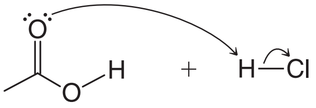
(a)



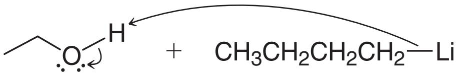
(b)



(c)



(d)



3.25 Write an equation, using the curved-arrow notation, for the acid–base reaction that will take place when each of the following are mixed. If no appreciable acid–base reaction takes place, because the equilibrium is unfavorable, you should so indicate.

(a) Aqueous NaOH and $CH_3CH_2CO_2H$

(b) Aqueous NaOH and $C_6H_5SO_3H$

(c) CH_3CH_2ONa in ethyl alcohol and ethyne

(d) CH_3CH_2Li in hexane and ethyne

(e) CH_3CH_2Li in hexane and ethyl alcohol

ACID–BASE STRENGTH AND EQUILIBRIA

3.26 When methyl alcohol is treated with NaH, the product is $CH_3O^-Na^+$ (and H_2) and not $Na^+\ ^-CH_2OH$ (and H_2). Explain why this is so.

3.27 What reaction will take place if ethyl alcohol is added to a solution of $HC{\equiv}C:^-\ Na^+$ in liquid ammonia?

3.28 **(a)** The K_a of formic acid (HCO_2H) is 1.77×10^{-4}. What is the pK_a? **(b)** What is the K_a of an acid whose $pK_a = 13$?

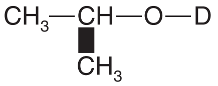
3.29 Acid HA has $pK_a = 20$; acid HB has $pK_a = 10$.

(a) Which is the stronger acid?

(b) Will an acid–base reaction with an equilibrium lying to the right take place if Na^+A^- is added to HB? Explain your answer.

3.30 Starting with appropriate unlabeled organic compounds, show syntheses of each of the following:

(a) $C_6H_5{-}C{\equiv}C{-}T$ **(b)**



$CH_3{-}\underset{CH_3}{\underset{|}{C}H}{-}O{-}D$

(c) $CH_3CH_2CH_2OD$

3.31 **(a)** Arrange the following compounds in order of decreasing acidity and explain your answer: $CH_3CH_2NH_2$, CH_3CH_2OH, and $CH_3CH_2CH_3$. **(b)** Arrange the conjugate bases of the acids given in part (a) in order of increasing basicity and explain your answer.

3.32 Arrange the following compounds in order of decreasing acidity:

(a) $CH_3CH{=}CH_2$, $CH_3CH_2CH_3$, $CH_3C{\equiv}CH$ **(b)** $CH_3CH_2CH_2OH$, $CH_3CH_2CO_2H$, $CH_3CHClCO_2H$

(c) CH_3CH_2OH, $CH_3CH_2OH_2^+$, CH_3OCH_3

3.33 Arrange the following in order of increasing basicity:

(a) CH_3NH_2, $CH_3NH_3^+$, CH_3NH^- **(b)** CH_3O^-, CH_3NH^-, $CH_3CH_2^-$

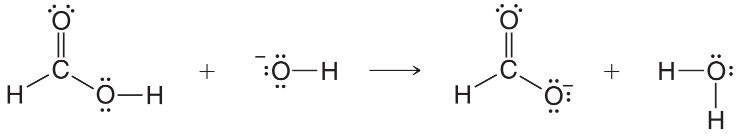
(c) $CH_3CH{=}CH^-$, $CH_3CH_2CH_2^-$, $CH_3C{\equiv}C^-$

GENERAL PROBLEMS

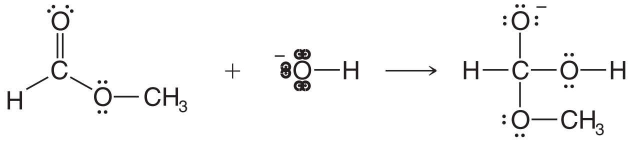
3.34 Whereas H_3PO_4 is a triprotic acid, H_3PO_3 is a diprotic acid. Draw structures for these two acids that account for this difference in behavior.

3.35 Supply the curved arrows necessary for the following reactions:

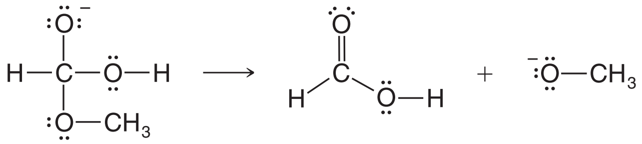
(a)



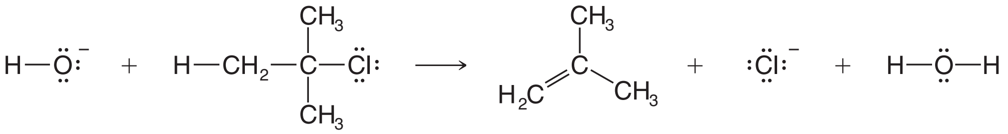
(b)



(c)



(d) $H-\ddot{\underset{..}{O}}:^- + CH_3-\ddot{\underset{..}{I}}: \longrightarrow H-\ddot{\underset{..}{O}}-CH_3 + :\ddot{\underset{..}{I}}:^-$

(e)


3.36 Glycine is an amino acid that can be obtained from most proteins. In solution, glycine exists in equilibrium between two forms:

$$H_2NCH_2CO_2H \rightleftharpoons H_3\overset{+}{N}CH_2CO_2^-$$

(a) Consult Table 3.1 and state which form is favored at equilibrium.

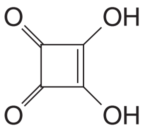
(b) A handbook gives the melting point of glycine as 262°C (with decomposition). Which of the structures given above best represents glycine?

3.37 Malonic acid, $HO_2CCH_2CO_2H$, is a diprotic acid. The pK_a for the loss of the first proton is 2.83; the pK_a for the loss of the second proton is 5.69. **(a)** Explain why malonic acid is a stronger acid than acetic acid (pK_a = 4.75). **(b)** Explain why the anion, $^-O_2CCH_2CO_2H$, is so much less acidic than malonic acid itself.

3.38 The free-energy change, $\Delta G°$, for the ionization of acid HA is 21 kJ mol^{-1}; for acid HB it is -21 kJ mol^{-1}. Which is the stronger acid?

3.39 At 25°C the enthalpy change, $\Delta H°$, for the ionization of trichloroacetic acid is $+6.3$ kJ mol^{-1} and the entropy change, $\Delta S°$, is $+0.0084$ kJ mol^{-1} K^{-1}. What is the pK_a of trichloroacetic acid?

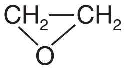
3.40 The compound at right has (for obvious reasons) been given the trivial name **squaric acid**. Squaric acid is a diprotic acid, with both protons being more acidic than acetic acid. In the dianion obtained after the loss of both protons, all of the carbon–carbon bonds are the same length as well as all of the carbon–oxygen bonds. Provide a resonance explanation for these observations.



Squaric acid

Challenge Problems

3.41 $CH_3CH_2SH + CH_3O^- \longrightarrow$ **A** (contains sulfur) + **B**

A +


$\longrightarrow$ **C** (which has the partial structure **A**—CH_2CH_2O)

C + $H_2O \longrightarrow$ **D** + **E** (which is inorganic)

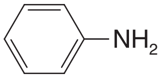
(a) Given the above sequence of reactions, draw structures for **A** through **E**.

(b) Rewrite the reaction sequence, showing all nonbonding electron pairs and using curved arrows to show electron pair movements.

3.42 First, complete and balance each of the equations below. Then, choosing among ethanol, hexane, and liquid ammonia, state which (there may be more than one) might be suitable solvents for each of these reactions. Disregard the practical limitations that come from consideration of "like dissolves like" and base your answers only on relative acidities.

(a) $CH_3(CH_2)_8OD + CH_3(CH_2)_8Li \longrightarrow$

(b) $NaNH_2 + CH_3C{\equiv}CH \longrightarrow$

(c) HCl +


$\longrightarrow$

(The conjugate acid of this amine, aniline, has a pK_a of 4.63.)

3.43 Dimethylformamide (DMF), $HCON(CH_3)_2$, is an example of a polar aprotic solvent, aprotic meaning it has no hydrogen atoms attached to highly electronegative atoms.

(a) Draw its dash-type structural formula, showing unshared electron pairs.

(b) Draw what you predict to be its most important resonance forms [one is your answer to part (a)].

(c) DMF, when used as the reaction solvent, greatly enhances the reactivity of nucleophiles (e.g., CN^- from sodium cyanide) in reactions like this:

$$NaCN + CH_3CH_2Br \longrightarrow CH_3CH_2C{\equiv}N + NaBr$$

Suggest an explanation for this effect of DMF on the basis of Lewis acid–base considerations. (*Hint*: Although water or an alcohol solvates both cations and anions, DMF is only effective in solvating cations.)

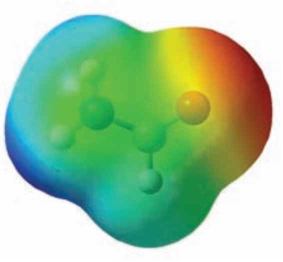
3.44 As noted in Table 3.1, the pK_a of acetone, CH_3COCH_3, is 19.2.

(a) Draw the bond-line formula of acetone and of any other contributing resonance form.

(b) Predict and draw the structure of the conjugate base of acetone and of any other contributing resonance form.

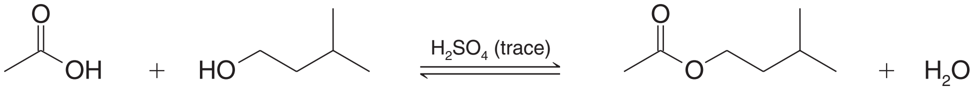
(c) Write an equation for a reaction that could be used to synthesize CH_3COCH_2D.

3.45 Formamide ($HCONH_2$) has a pK_a of approximately 25. Predict, based on the map of electrostatic potential for formamide shown here, which hydrogen atom(s) has this pK_a value. Support your conclusion with arguments having to do with the electronic structure of formamide.



Learning Group Problems

Suppose you carried out the following synthesis of 3-methylbutyl ethanoate (isoamyl acetate):



Ethanoic acid (excess) **3-Methyl-1-butanol** **3-Methylbutyl ethanoate**

As the chemical equation shows, 3-methyl-1-butanol (also called isoamyl alcohol or isopentyl alcohol) was mixed with an excess of acetic acid (ethanoic acid by its systematic name) and a trace of sulfuric acid (which serves as a catalyst). This reaction is an equilibrium reaction, so it is expected that not all of the starting materials will be consumed. The equilibrium should lie quite far to the right due to the excess of acetic acid used, but not completely.

After an appropriate length of time, isolation of the desired product from the reaction mixture was begun by adding a volume of 5% aqueous sodium bicarbonate ($NaHCO_3$ has an effective pK_a of 7) roughly equal to the volume of the reaction mixture. Bubbling occurred and a mixture consisting of two layers resulted—a basic aqueous layer and an organic layer. The layers were separated and the aqueous layer was removed. The addition of aqueous sodium bicarbonate to the layer of organic materials and separation of the layers were repeated twice. Each time the predominantly aqueous layers were removed, they were combined in the same collection flask. The organic layer that remained after the three bicarbonate extractions was dried and then subjected to distillation in order to obtain a pure sample of 3-methylbutyl ethanoate (isoamyl acetate).

1. List all the chemical species likely to be present at the end of the reaction but before adding aqueous $NaHCO_3$. Note that the H_2SO_4 was not consumed (since it is a catalyst), and is thus still available to donate a proton to atoms that can be protonated.

2. Use a table of pK_a values, such as Table 3.1, to estimate pK_a values for any potentially acidic hydrogens in each of the species you listed in part 1 (or for the conjugate acid).

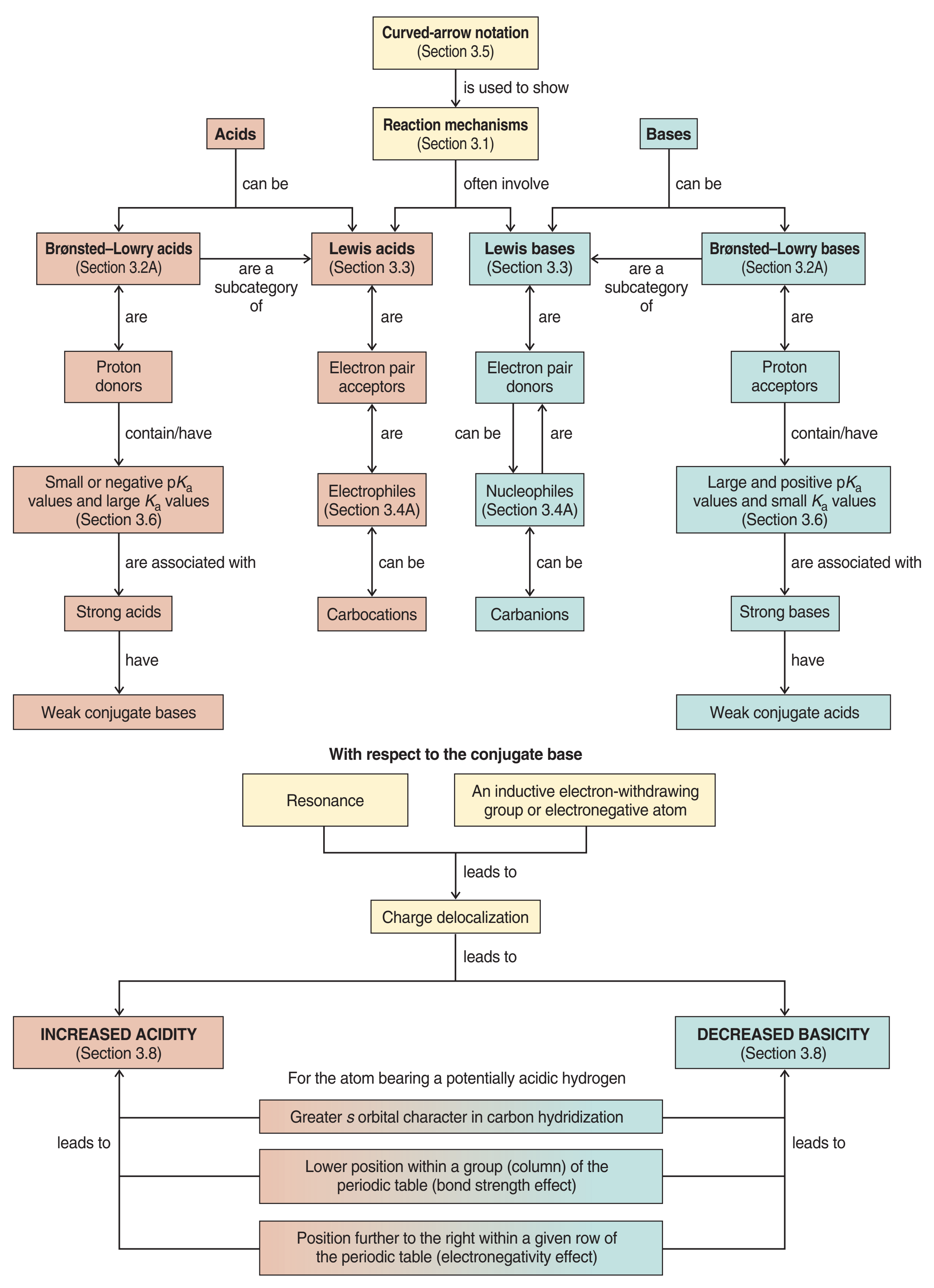
3. Write chemical equations for all the acid–base reactions you would predict to occur (based on the pK_a values you used) when the species you listed above encounter the aqueous sodium bicarbonate solution. (*Hint*: Consider whether each species might be an acid that could react with $NaHCO_3$.)

4. **(a)** Explain, on the basis of polarities and solubility, why separate layers formed when aqueous sodium bicarbonate was added to the reaction mixture. (*Hint*: Most sodium salts of organic acids are soluble in water, as are neutral oxygen-containing organic compounds of four carbons or less.)

(b) List the chemical species likely to be present after the reaction with $NaHCO_3$ in (i) the organic layer and (ii) the aqueous layer.

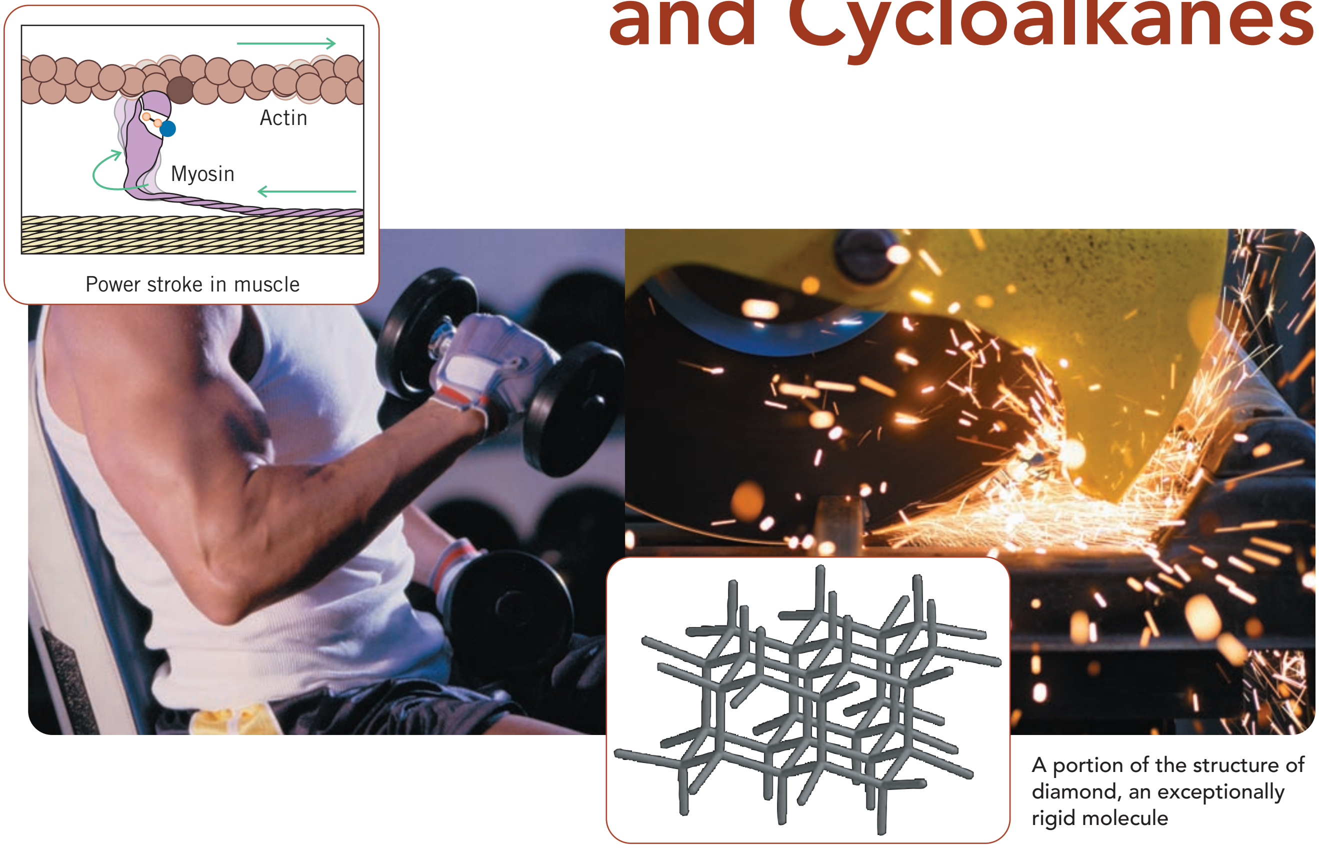
(c) Why was the aqueous sodium bicarbonate extraction step repeated three times?

CONCEPT MAP


Curved-arrow notation (Section 3.5)
is used to show
Reaction mechanisms (Section 3.1)
often involve
Acids
can be
Bases
can be
Brønsted–Lowry acids (Section 3.2A)
are a subcategory of
Lewis acids (Section 3.3)
Lewis bases (Section 3.3)
are a subcategory of
Brønsted–Lowry bases (Section 3.2A)
are
are
are
are
Proton donors
Electron pair acceptors
Electron pair donors
Proton acceptors
contain/have
are
can be
are
contain/have
Small or negative pK_a values and large K_a values (Section 3.6)
Electrophiles (Section 3.4A)
Nucleophiles (Section 3.4A)
Large and positive pK_a values and small K_a values (Section 3.6)
are associated with
can be
can be
are associated with
Strong acids
Carbocations
Carbanions
Strong bases
have
have
Weak conjugate bases
Weak conjugate acids
With respect to the conjugate base
Resonance
An inductive electron-withdrawing group or electronegative atom
leads to
Charge delocalization
leads to
INCREASED ACIDITY (Section 3.8)
DECREASED BASICITY (Section 3.8)
For the atom bearing a potentially acidic hydrogen
Greater *s* orbital character in carbon hydridization
Lower position within a group (column) of the periodic table (bond strength effect)
Position further to the right within a given row of the periodic table (electronegativity effect)
leads to
leads to

4

Nomenclature and Conformations of Alkanes and Cycloalkanes

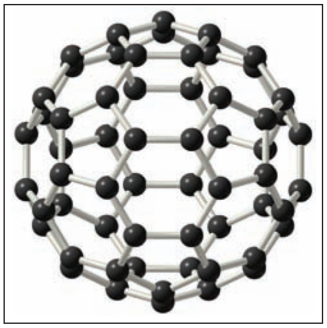


Power stroke in muscle

A portion of the structure of diamond, an exceptionally rigid molecule

When your muscles contract to do work, like the person shown exercising above, it is largely because many carbon–carbon sigma (single) bonds are undergoing rotation (conformational changes) in a muscle protein called myosin. But when a diamond-tipped blade cuts, as shown above, the carbon–carbon single bonds comprising the diamond resist all the forces brought to bear on them, such that the material yields to the diamond. This remarkable contrast in properties, from the flexibility of muscles to the rigidity of diamond, depends on many things, but central to them is whether or not rotation is possible around carbon–carbon bonds. In this chapter we shall consider such bond rotations.

We learned in Chapter 2 that our study of organic chemistry can be organized around functional groups. Now we consider the hydrocarbon framework to which functional groups are attached—the framework that consists of only carbon and hydrogen atoms. From the standpoint of an architect, hydrocarbon frameworks present a dream of limitless possibilities, which is part of what makes organic chemistry such a fascinating discipline. Buckminsterfullerene, named after the visionary architect Buckminster Fuller, is just one example of a hydrocarbon with intriguing molecular architecture.



Buckminsterfullerene

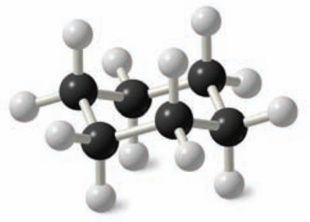
Polyethylene is a hydrocarbon macromolecule that is inert to most chemicals we use in day-to-day life.

Even though there are vast possibilities for the structures of organic molecules, fortunately there is a well-defined system for naming carbon molecules. We study the essentials of this system here in Chapter 4, and then build on it as we study the chemistry of functional groups in later chapters.

When chemists talk about structure in organic chemistry, they mean not only the connectivity of the atoms, but also the shapes that molecules can adopt due to rotations of groups joined by single bonds. The examination of these properties is called conformational analysis, which we also discuss in this chapter as we consider the carbon framework of organic molecules.

We also consider the properties and reactivity of hydrocarbons. Under ambient conditions, hydrocarbons containing only carbon–carbon single bonds are relatively inert. Polyethylene, for example, is a hydrocarbon that is used for household containers, tubing, and many other items where lack of reactivity is important. Hydrocarbons are combustible, however, and of course we make use of this property every time we burn hydrocarbon fuels such as natural gas, gasoline, or diesel. The release of greenhouse gases by combustion of hydrocarbons is a concern regarding climate change, of course.

4.1 *Introduction to Alkanes and Cycloalkanes*



Cyclohexane

We noted earlier that the family of organic compounds called hydrocarbons can be divided into several groups on the basis of the type of bond that exists between the individual carbon atoms. Those hydrocarbons in which all of the carbon–carbon bonds are single bonds are called **alkanes**, those hydrocarbons that contain a carbon–carbon double bond are called **alkenes**, and those with a carbon–carbon triple bond are called **alkynes**.

Cycloalkanes are alkanes in which all or some of the carbon atoms are arranged in a ring. Alkanes have the general formula C_nH_{2n+2}; cycloalkanes containing a single ring have two fewer hydrogen atoms and thus have the general formula C_nH_{2n}.

Alkanes and cycloalkanes are so similar that many of their properties can be considered side by side. Some differences remain, however, and certain structural features arise from the rings of cycloalkanes that are more conveniently studied separately. We shall point out the chemical and physical similarities of alkanes and cycloalkanes as we go along.

4.1A Sources of Alkanes: Petroleum

The primary source of alkanes is petroleum. Petroleum is a complex mixture of organic compounds, most of which are alkanes and aromatic hydrocarbons (cf. Chapter 14). It also contains small amounts of oxygen-, nitrogen-, and sulfur-containing compounds.

Some of the molecules in petroleum are clearly of biological origin. The natural origin of petroleum is still under debate, however. Many scientists believe petroleum originated with decay of primordial biological matter. Recent theories suggest, however, that organic molecules may have been included as Earth formed by accretion of interstellar materials. Analysis of asteroids and comets has shown that they contain a significant amount and variety of organic compounds. Methane and other hydrocarbons are found in the atmospheres of Jupiter, Saturn, and Uranus. Saturn's moon Titan has a solid form of methane–water ice at its surface and an atmosphere rich in methane. Earth's petroleum may therefore have originated similarly to the way methane became part of these other bodies in our solar system. The discovery of microbial life in high-temperature ocean vents and the growing evidence for a deep, hot biosphere within Earth suggest that compounds in petroleum of biological origin may simply be contaminants introduced by primitive life into a nonbiologically formed petroleum reserve that was present from Earth's beginning.

Petroleum is a finite resource whose origin is under debate. At the La Brea Tar Pits in Los Angeles, many prehistoric animals perished in a natural vat containing hydrocarbons.

THE CHEMISTRY OF . . .

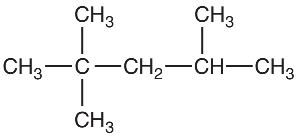
Petroleum Refining

A petroleum refinery. The tall towers are fractioning columns used to separate components of crude oil according to their boiling points.

The first step in refining petroleum is distillation; the object here is to separate the petroleum into fractions based on the volatility of its components. Complete separation into fractions containing individual compounds is economically impractical and virtually impossible technically. More than 500 different compounds are contained in the petroleum distillates boiling below 200°C, and many have almost the same boiling points. Thus the fractions taken contain mixtures of alkanes of similar boiling points (see the table below). Mixtures of alkanes, fortunately, are perfectly suitable for uses as fuels, solvents, and lubricants, the primary uses of petroleum.

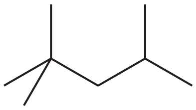
The demand for gasoline is much greater than that supplied by the gasoline fraction of petroleum. Important processes in the petroleum industry, therefore, are concerned with converting hydrocarbons from other fractions into gasoline. When a mixture of alkanes from the gas oil fraction (C_{12} and higher) is heated at very high temperatures (~500°C) in the presence of a variety of catalysts, the molecules break apart and rearrange to smaller, more highly branched hydrocarbons containing 5–10 carbon atoms. This process is called *catalytic cracking*. **Cracking** can also be done in the absence of a catalyst—called ***thermal cracking***—but in this process the products tend to have unbranched chains, and alkanes with unbranched chains have a very low "octane rating."

The highly branched compound 2,2,4-trimethylpentane (called isooctane in the petroleum industry) burns very smoothly (without knocking) in internal combustion engines and is used as one of the standards by which the octane rating of gasolines is established.

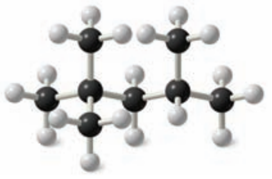


$$\mathrm{CH_3-\overset{\overset{\displaystyle CH_3}{|}}{\underset{\underset{\displaystyle CH_3}{|}}{C}}-CH_2-\overset{\overset{\displaystyle CH_3}{|}}{CH}-CH_3}$$

or



or



2,2,4-Trimethylpentane ("isooctane")

According to this scale, 2,2,4-trimethylpentane has an octane rating of 100. Heptane, $CH_3(CH_2)_5CH_3$, a compound that produces much knocking when it is burned in an internal combustion engine, is given an octane rating of 0. Mixtures of 2,2,4-trimethylpentane and heptane are used as standards for octane ratings between 0 and 100. A gasoline, for example, that has the same characteristics in an engine as a mixture of 87% 2,2,4-trimethylpentane and 13% heptane would be rated as 87-octane gasoline.

Typical Fractions Obtained by Distillation of Petroleum

Boiling Range of Fraction (°C)	Number of Carbon Atoms per Molecule	Use
Below 20	C_1–C_4	Natural gas, bottled gas, petrochemicals
20–60	C_5–C_6	Petroleum ether, solvents
60–100	C_6–C_7	Ligroin, solvents
40–200	C_5–C_{10}	Gasoline (straight-run gasoline)
175–325	C_{12}–C_{18}	Kerosene and jet fuel
250–400	C_{12} and higher	Gas oil, fuel oil, and diesel oil
Nonvolatile liquids	C_{20} and higher	Refined mineral oil, lubricating oil, and grease
Nonvolatile solids	C_{20} and higher	Paraffin wax, asphalt, and tar

Adapted with permission of John Wiley & Sons, Inc., from Holum, J. R., *General, Organic, and Biological Chemistry, Ninth Edition*, p. 213. Copyright 1995.

4.2 *Shapes of Alkanes*

A general tetrahedral orientation of groups—and thus sp^3 hybridization—is the rule for the carbon atoms of all alkanes and cycloalkanes. We can represent the shapes of alkanes as shown in Fig. 4.1.

Butane and pentane are examples of alkanes that are sometimes called "straight-chain" alkanes. One glance at three-dimensional models, however, shows that because of their tetrahedral carbon atoms the chains are zigzagged and not at all straight. Indeed, the structures that we have depicted in Fig. 4.1 are the straightest possible arrangements of the chains because rotations about the carbon–carbon single bonds produce arrangements that are even less straight. A better description is **unbranched**. This means that each carbon atom within the chain is bonded to no more than two other carbon atoms and that unbranched alkanes contain only primary and secondary carbon atoms. Primary, secondary, and tertiary carbon atoms were defined in Section 2.5.

Isobutane, isopentane, and neopentane (Fig. 4.2) are examples of branched-chain alkanes. In neopentane the central carbon atom is bonded to four carbon atoms.

Butane and isobutane have the same molecular formula: C_4H_{10}. The two compounds have their atoms connected in a different order and are, therefore, **constitutional isomers** (Section 1.3). Pentane, isopentane, and neopentane are also constitutional isomers. They, too, have the same molecular formula (C_5H_{12}) but have different structures.

Figure 4.1 Ball-and-stick models for three simple alkanes.

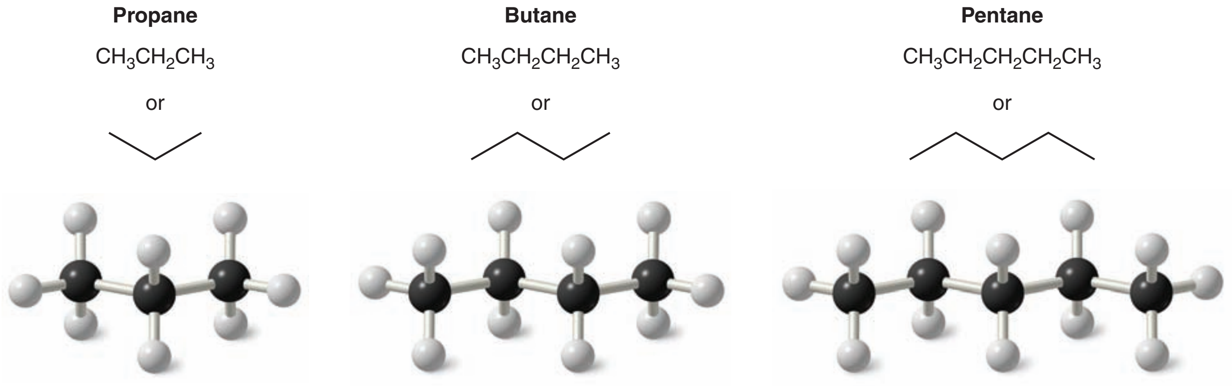
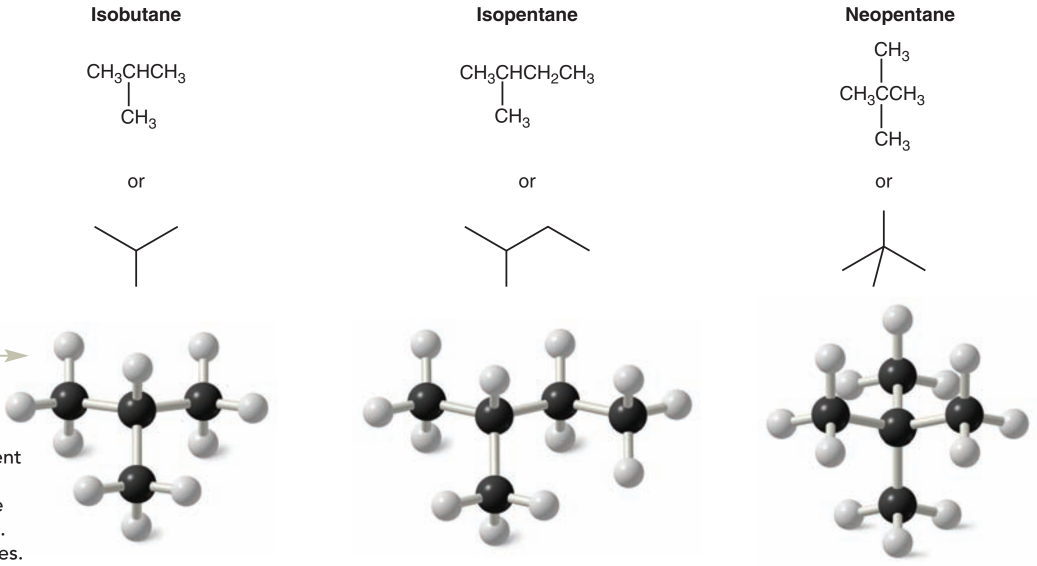


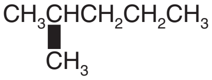
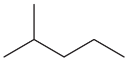
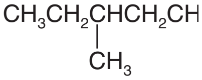
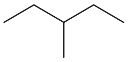
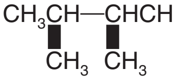
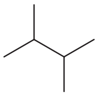
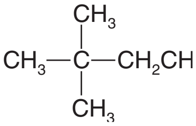
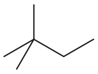
Figure 4.2 Ball-and-stick models for three branched-chain alkanes. In each of the compounds one carbon atom is attached to more than two other carbon atoms.



Helpful Hint

You should build your own molecular models of the compounds in Figs. 4.1 and 4.2. View them from different perspectives and experiment with how their shapes change when you twist various bonds. Make drawings of your structures.

TABLE 4.1 Physical Constants of the Hexane Isomers

Molecular Formula	Condensed Structural Formula	Bond-Line Formula	mp (°C)	bp (°C)[a] (1 atm)	Density[b] (g mL^{-1})	Index of Refraction[c] (n_D 20°C)
C_6H_{14}	$CH_3CH_2CH_2CH_2CH_2CH_3$		−95	68.7	0.6594^{20}	1.3748
C_6H_{14}	 $CH_3\underset{CH_3}{CH}CH_2CH_2CH_3$		−153.7	60.3	0.6532^{20}	1.3714
C_6H_{14}	 $CH_3CH_2\underset{CH_3}{CH}CH_2CH$		−118	63.3	0.6643^{20}	1.3765
C_6H_{14}	 $CH_3\underset{CH_3}{CH}—\underset{CH_3}{CH}CH$		−128.8	58	0.6616^{20}	1.3750
C_6H_{14}	 $CH_3—\overset{CH_3}{\underset{CH_3}{C}}—CH_2CH$		−98	49.7	0.6492^{20}	1.3688

[a] Unless otherwise indicated, all boiling points given in this book are at 1 atm or 760 torr.

[b] The superscript indicates the temperature at which the density was measured.

[c] The index of refraction is a measure of the ability of the alkane to bend (refract) light rays. The values reported are for light of the D line of the sodium spectrum (n_D).

Review Problem 4.1

Write condensed and bond-line structural formulas for all of the constitutional isomers with the molecular formula C_7H_{16}. (There is a total of nine constitutional isomers.)

Constitutional isomers, as stated earlier, have different physical properties. The differences may not always be large, but constitutional isomers are always found to have different melting points, boiling points, densities, indexes of refraction, and so forth. Table 4.1 gives some of the physical properties of the C_6H_{14} isomers.

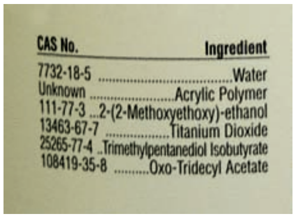
As Table 4.2 shows, the number of constitutional isomers that is possible increases dramatically as the number of carbon atoms in the alkane increases.

The large numbers in Table 4.2 are based on calculations that must be done with a computer. Similar calculations, which take into account stereoisomers (Chapter 5) as well as constitutional isomers, indicate that an alkane with the formula $C_{167}H_{336}$ would, in theory, have more possible isomers than there are particles in the observed universe!

TABLE 4.2 Number of Alkane Isomers

Molecular Formula	Possible Number of Constitutional Isomers
C_4H_{10}	2
C_5H_{12}	3
C_6H_{14}	5
C_7H_{16}	9
C_8H_{18}	18
C_9H_{20}	35
$C_{10}H_{22}$	75
$C_{15}H_{32}$	4,347
$C_{20}H_{42}$	366,319
$C_{30}H_{62}$	4,111,846,763
$C_{40}H_{82}$	62,481,801,147,341

4.3 *IUPAC Nomenclature of Alkanes, Alkyl Halides, and Alcohols*



CAS No.	Ingredient
7732-18-5	Water
Unknown	Acrylic Polymer
111-77-3	2-(2-Methoxyethoxy)-ethanol
13463-67-7	Titanium Dioxide
25265-77-4	Trimethylpentanediol Isobutyrate
108419-35-8	Oxo-Tridecyl Acetate

The Chemical Abstracts Service assigns a CAS Registry Number to every compound. CAS numbers make it easy to find information about a compound in the chemical literature. The CAS numbers for ingredients in a can of latex paint are shown here.

Prior to the development near the end of the nineteenth century of a formal system for naming organic compounds, many organic compounds had already been discovered or synthesized. Early chemists named these compounds, often on the basis of the source of the compound. Acetic acid (systematically called ethanoic acid) is an example; it was obtained by distilling vinegar, and it got its name from the Latin word for vinegar, *acetum*. Formic acid (systematically called methanoic acid) had been obtained by the distillation of the bodies of ants, so it got the name from the Latin word for ants, *formicae*. Many of these older names for compounds, called common or trivial names, are still in wide use today.

Today, chemists use a systematic nomenclature developed and updated by the International Union of Pure and Applied Chemistry (IUPAC). Underlying the IUPAC system is a fundamental principle: **each different compound should have a different and unambiguous name.***

The **IUPAC system** for naming alkanes is not difficult to learn, and the principles involved are used in naming compounds in other families as well. For these reasons we begin our study of the IUPAC system with the rules for naming alkanes and then study the rules for alkyl halides and alcohols.

The names for several of the unbranched alkanes are listed in Table 4.3. The ending for all of the names of alkanes is *-ane*. The stems of the names of most of the alkanes (above C_4) are of Greek and Latin origin. Learning the stems is like learning to count in organic chemistry. Thus, one, two, three, four, and five become meth-, eth-, prop-, but-, and pent-.

TABLE 4.3 The Unbranched Alkanes

Name	Number of Carbon Atoms	Structure	Name	Number of Carbon Atoms	Structure
Methane	1	CH_4	Undecane	11	$CH_3(CH_2)_9CH_3$
Ethane	2	CH_3CH_3	Dodecane	12	$CH_3(CH_2)_{10}CH_3$
Propane	3	$CH_3CH_2CH_3$	Tridecane	13	$CH_3(CH_2)_{11}CH_3$
Butane	4	$CH_3(CH_2)_2CH_3$	Tetradecane	14	$CH_3(CH_2)_{12}CH_3$
Pentane	5	$CH_3(CH_2)_3CH_3$	Pentadecane	15	$CH_3(CH_2)_{13}CH_3$
Hexane	6	$CH_3(CH_2)_4CH_3$	Hexadecane	16	$CH_3(CH_2)_{14}CH_3$
Heptane	7	$CH_3(CH_2)_5CH_3$	Heptadecane	17	$CH_3(CH_2)_{15}CH_3$
Octane	8	$CH_3(CH_2)_6CH_3$	Octadecane	18	$CH_3(CH_2)_{16}CH_3$
Nonane	9	$CH_3(CH_2)_7CH_3$	Nonadecane	19	$CH_3(CH_2)_{17}CH_3$
Decane	10	$CH_3(CH_2)_8CH_3$	Eicosane	20	$CH_3(CH_2)_{18}CH_3$

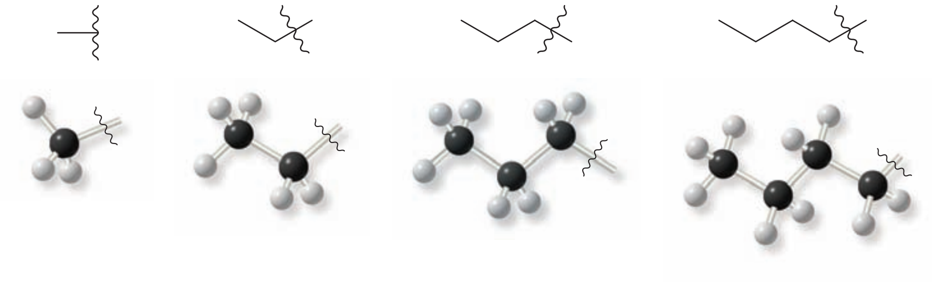
4.3A Nomenclature of Unbranched Alkyl Groups

If we remove one hydrogen atom from an alkane, we obtain what is called an **alkyl group**. These alkyl groups have names that end in **-yl**. When the alkane is **unbranched**, and the hydrogen atom that is removed is a **terminal** hydrogen atom, the names are straightforward:

---

*The complete IUPAC rules for nomenclature can be found through links at the IUPAC website (www.iupac.org).

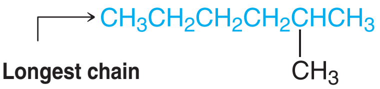
CH_3—**H**	CH_3CH_2—**H**	$CH_3CH_2CH_2$—**H**	$CH_3CH_2CH_2CH_2$—**H**
Methane	**Ethane**	**Propane**	**Butane**
CH_3—	CH_3CH_2—	$CH_3CH_2CH_2$—	$CH_3CH_2CH_2CH_2$—
Methyl	**Ethyl**	**Propyl**	**Butyl**
Me-	**Et-**	**Pr-**	**Bu-**



4.3B Nomenclature of Branched-Chain Alkanes

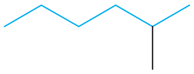
Branched-chain alkanes are named according to the following rules:

1. **Locate the longest continuous chain of carbon atoms; this chain determines the parent name for the alkane**. We designate the following compound, for example, as a *hexane* because the longest continuous chain contains six carbon atoms:

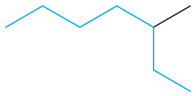


$$CH_3CH_2CH_2CH_2\underset{\displaystyle CH_3}{\underset{|}{C}}HCH_3$$

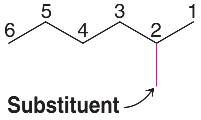
or

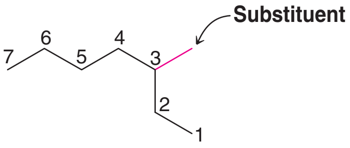


The longest continuous chain may not always be obvious from the way the formula is written. Notice, for example, that the following alkane is designated as a *heptane* because the longest chain contains seven carbon atoms:

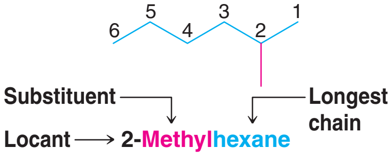


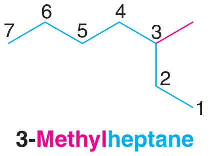
2. **Number the longest chain beginning with the end of the chain nearer the substituent**. Applying this rule, we number the two alkanes that we illustrated previously in the following way:





3. **Use the numbers obtained by application of rule 2 to designate the location of the substituent group**. The parent name is placed last, and the substituent group, preceded by the number designating its location on the chain, is placed first. Numbers are separated from words by a hyphen. Our two examples are 2-methylhexane and 3-methylheptane, respectively:



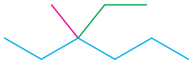


4. **When two or more substituents are present, give each substituent a number corresponding to its location on the longest chain**. For example, we designate the following compound as 4-ethyl-2-methylhexane:

4-Ethyl-2-methylhexane

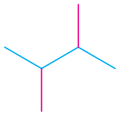
The substituent groups should be listed *alphabetically* (i.e., ethyl before methyl).* In deciding on alphabetical order, disregard multiplying prefixes such as "di" and "tri."

5. **When two substituents are present on the same carbon atom, use that number twice**:

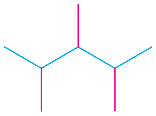


3-Ethyl-3-methylhexane

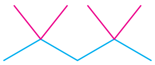
6. **When two or more substituents are identical, indicate this by the use of the prefixes di-, tri-, tetra-**, and so on. Then make certain that each and every substituent has a number. Commas are used to separate numbers from each other:



2,3-Dimethylbutane



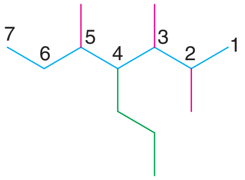
2,3,4-Trimethylpentane



2,2,4,4-Tetramethylpentane

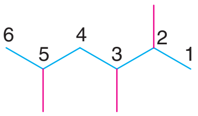
Application of these six rules allows us to name most of the alkanes that we shall encounter. Two other rules, however, may be required occasionally:

7. **When two chains of equal length compete for selection as the parent chain, choose the chain with the greater number of substituents**:



2,3,5-Trimethyl-4-propylheptane
(four substituents)

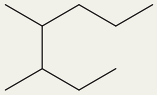
8. **When branching first occurs at an equal distance from either end of the longest chain, choose the name that gives the lower number at the first point of difference**:



2,3,5-Trimethylhexane
(*not* 2,4,5-trimethylhexane)

Solved Problem 4.1

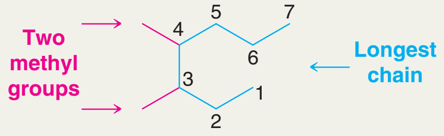
Provide an IUPAC name for the following alkane.



---

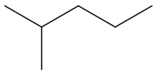
*Some handbooks also list the groups in order of increasing size or complexity (i.e., methyl before ethyl). An alphabetical listing, however, is now by far the most widely used system.

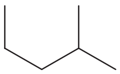
STRATEGY AND SOLUTION We find the longest chain (shown in blue) to be seven carbons; therefore the parent name is heptane. There are two methyl substituents (shown in red). We number the chain so as to give the first methyl group the lower number. The correct name, therefore, is 3,4-dimethylheptane. Numbering the chain from the other end to give 4,5-dimethylheptane would have been incorrect.

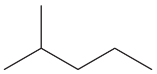


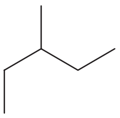
Review Problem 4.2

Which structure does not represent 2-methylpentane?

(a)


(b)


(c)


(d)


Review Problem 4.3

Write the structure and give the IUPAC name for an alkane with formula C_6H_{14} that has only primary and secondary carbon atoms.

Review Problem 4.4

Draw bond-line formulas for all of the isomers of C_8H_{18} that have (a) methyl substituents, and (b) ethyl substituents.

4.3C Nomenclature of Branched Alkyl Groups

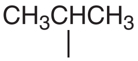
In Section 4.3A you learned the names for the unbranched alkyl groups such as methyl, ethyl, propyl, and butyl, groups derived by removing a terminal hydrogen from an alkane. For alkanes with more than two carbon atoms, more than one derived group is possible. Two groups can be derived from propane, for example; the **propyl group** is derived by removal of a terminal hydrogen, and the **1-methylethyl** or **isopropyl group** is derived by removal of a hydrogen from the central carbon:

Three-Carbon Groups

$CH_3CH_2CH_3$

Propane

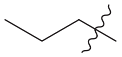
$CH_3CH_2CH_2—$

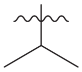


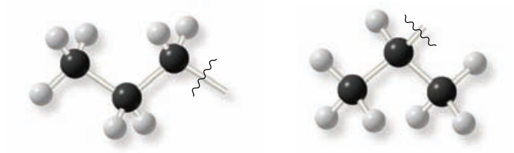
$CH_3\underset{|}{C}HCH_3$

Propyl **Isopropyl**

Pr- i-Pr-







1-Methylethyl is the systematic name for this group; isopropyl is a common name. Systematic nomenclature for alkyl groups is similar to that for branched-chain alkanes, with the provision that *numbering always begins at the point where the group is attached to the main chain*. There are four C_4 groups.

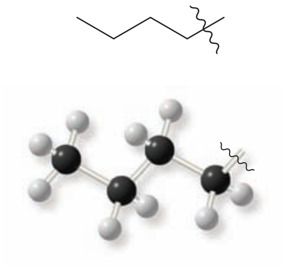
Four-Carbon Groups

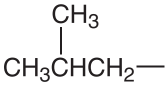
$CH_3CH_2CH_2CH_3$

Butane

$CH_3CH_2CH_2CH_2-$

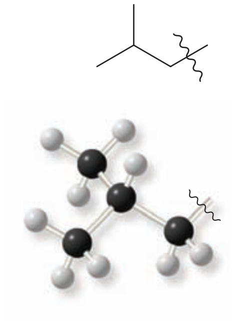
Butyl

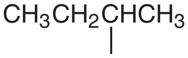




$$\begin{array}{l} \quad\;\; CH_3 \\ \quad\;\;\, | \\ CH_3CHCH_2- \end{array}$$

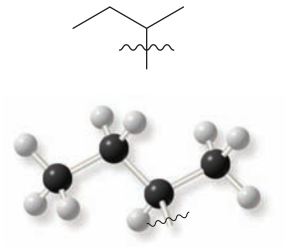
Isobutyl





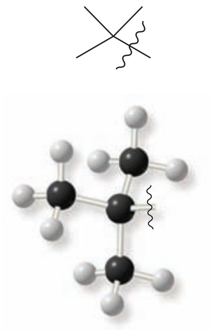
$$\begin{array}{l} CH_3CH_2CHCH_3 \\ \qquad\quad\;\;\, | \end{array}$$

***sec*-Butyl**

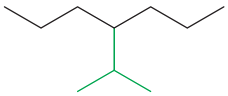


$(CH_3)_3C-$

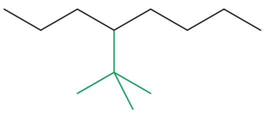
***tert*-Butyl (or *t*-Bu)**



The following examples show how the names of these groups are employed:



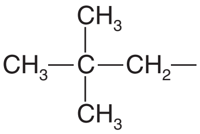
4-(1-Methylethyl)heptane or 4-isopropylheptane



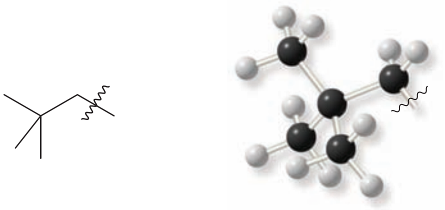
4-(1,1-Dimethylethyl)octane or 4-*tert*-butyloctane

The common names **isopropyl, isobutyl**, ***sec*-butyl**, and ***tert*-butyl** are approved by the IUPAC for the unsubstituted groups, and they are still very frequently used. You should learn these groups so well that you can recognize them any way that they are written. In deciding on alphabetical order for these groups you should disregard structure-defining prefixes that are written in italics and separated from the name by a hyphen. Thus *tert*-butyl precedes ethyl, but ethyl precedes isobutyl.*

There is one five-carbon group with an IUPAC approved common name that you should also know: the 2,2-dimethylpropyl group, commonly called the **neopentyl group**:



$$\begin{array}{c} CH_3 \\ | \\ CH_3-C-CH_2- \\ | \\ CH_3 \end{array}$$



2,2-Dimethylpropyl or neopentyl group

Review Problem 4.5

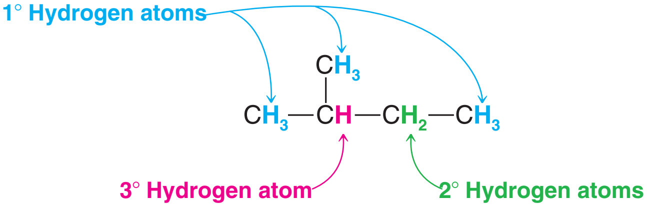
(a) In addition to the 2,2-dimethylpropyl (or neopentyl) group just given, there are seven other five-carbon groups. Draw bond-line formulas for their structures and give each structure its systematic name. **(b)** Draw bond-line formulas and provide IUPAC names for all of the isomers of C_7H_{16}.

---

*The abbreviations *i*, *s*, and *t* are sometimes used for iso-, *sec*-, and *tert*-, respectively.

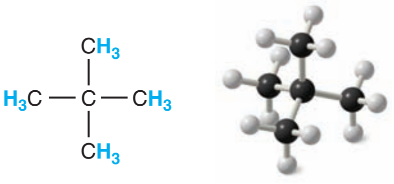
4.3D Classification of Hydrogen Atoms

The hydrogen atoms of an alkane are classified on the basis of the carbon atom to which they are attached. A hydrogen atom attached to a primary carbon atom is a primary (1°) hydrogen atom, and so forth. The following compound, 2-methylbutane, has primary, secondary (2°), and tertiary (3°) hydrogen atoms:



$$CH_3-\overset{\displaystyle CH_3}{CH}-CH_2-CH_3$$

On the other hand, 2,2-dimethylpropane, a compound that is often called **neopentane**, has only primary hydrogen atoms:



$$H_3C-\overset{\displaystyle CH_3}{\underset{\displaystyle CH_3}{C}}-CH_3$$

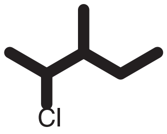
2,2-Dimethylpropane (neopentane)

4.3E Nomenclature of Alkyl Halides

Alkanes bearing halogen substituents are named in the IUPAC substitutive system as haloalkanes:

CH_3CH_2Cl	$CH_3CH_2CH_2F$	$CH_3CHBrCH_3$
Chloroethane	**1-Fluoropropane**	**2-Bromopropane**

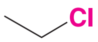
- When the parent chain has both a halo and an alkyl substituent attached to it, number the chain from the end nearer the first substituent, regardless of whether it is halo or alkyl. If two substituents are at equal distance from the end of the chain, then number the chain from the end nearer the substituent that has alphabetical precedence:



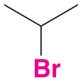
2-Chloro-3-methylpentane

2-Chloro-4-methylpentane

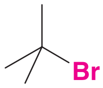
Common names for many simple haloalkanes are still widely used, however. In this common nomenclature system, called **functional class nomenclature**, haloalkanes are named as alkyl halides. (The following names are also accepted by the IUPAC.)



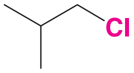
Ethyl chloride



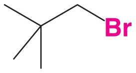
Isopropyl bromide



***tert*-Butyl bromide**



Isobutyl chloride



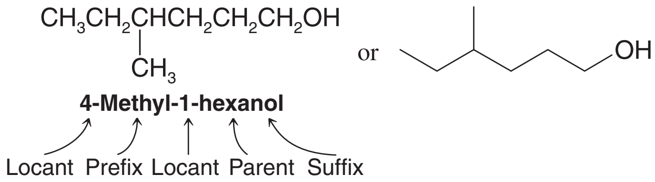
Neopentyl bromide

Review Problem 4.6

Draw bond-line formulas and give IUPAC substitutive names for all of the isomers of **(a)** C_4H_9Cl and **(b)** $C_5H_{11}Br$.

4.3F Nomenclature of Alcohols

In what is called IUPAC **substitutive nomenclature**, a name may have as many as four features: **locants, prefixes, parent compound**, and **suffixes**. Consider the following compound as an illustration without, for the moment, being concerned as to how the name arises:



$$\underset{\displaystyle CH_3}{CH_3CH_2\overset{|}{C}HCH_2CH_2CH_2OH}$$

4-Methyl-1-hexanol

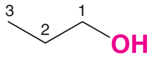
The *locant* **4-** tells that the substituent **methyl** group, named as a *prefix*, is attached to the *parent compound* at C4. The parent compound contains six carbon atoms and no multiple bonds, hence the parent name **hexane**, and it is an alcohol; therefore it has the *suffix* **-ol**. The locant **1-** tells that C1 bears the hydroxyl group. **In general, numbering of the chain always begins at the end nearer the group named as a suffix**.

The locant for a suffix (whether it is for an alcohol or another functional group) may be placed before the parent name as in the above example or, according to a 1993 IUPAC revision of the rules, immediately before the suffix. Both methods are IUPAC approved. Therefore, the above compound could also be named **4-methylhexan-1-ol**.

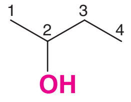
The following procedure should be followed in giving alcohols IUPAC substitutive names:

1. Select the longest continuous carbon chain *to which the hydroxyl is directly attached*. Change the name of the alkane corresponding to this chain by dropping the final *-e* and adding the suffix *-ol*.
2. Number the longest continuous carbon chain so as to give the carbon atom bearing the hydroxyl group the lower number. Indicate the position of the hydroxyl group by using this number as a locant; indicate the positions of other substituents (as prefixes) by using the numbers corresponding to their positions along the carbon chain as locants.

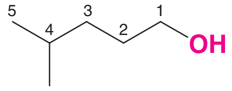
The following examples show how these rules are applied:



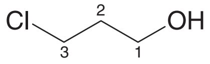
1-Propanol



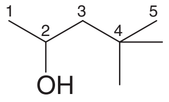
2-Butanol



4-Methyl-1-pentanol
or 4-methylpentan-1-ol
(*not* 2-methyl-5-pentanol)



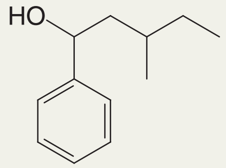
3-Chloro-1-propanol
or 3-chloropropan-1-ol



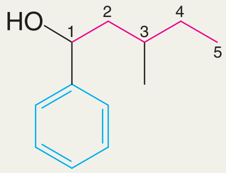
4,4-Dimethyl-2-pentanol
or 4,4-dimethylpentan-2-ol

Solved Problem 4.2

Give an IUPAC name for the following compound.



STRATEGY AND ANSWER We find that the longest carbon chain (in red at right) has five carbons and it bears a hydroxyl group on the first carbon. So we name this part of the molecule as a 1-pentanol. There is a phenyl group on carbon-1 and a methyl group on carbon-3, so the full name is 3-methyl-1-phenyl-1-pentanol.



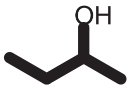
Review Problem 4.7

Draw bond-line formulas and give IUPAC substitutive names for all of the isomeric alcohols with the formulas **(a)** $C_4H_{10}O$ and **(b)** $C_5H_{12}O$.

Simple alcohols are often called by *common* functional class names that are also approved by the IUPAC. We have seen several examples already (Section 2.6). In addition to *methyl alcohol, ethyl alcohol*, and *isopropyl alcohol*, there are several others, including the following:

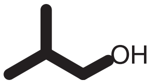
Propyl alcohol

Butyl alcohol



***sec*-Butyl alcohol**

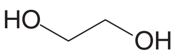
***tert*-Butyl alcohol**

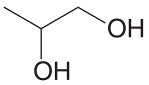


Isobutyl alcohol

Neopentyl alcohol

Alcohols containing two hydroxyl groups are commonly called **glycols**. In the IUPAC substitutive system they are named as **diols**:



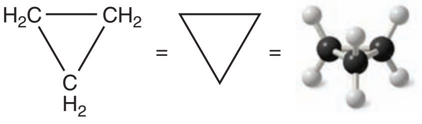


Substitutive	**1,2-Ethanediol or ethane-1,2-diol**	**1,2-Propanediol or propane-1,2-diol**	**1,3-Propanediol or propane-1,3-diol**
Common	***Ethylene glycol***	***Propylene glycol***	***Trimethylene glycol***

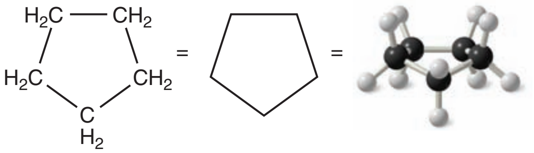
4.4 How to Name Cycloalkanes

4.4A Monocyclic Compounds

Cycloalkanes with only one ring are named by attaching the prefix cyclo- to the names of the alkanes possessing the same number of carbon atoms. For example,



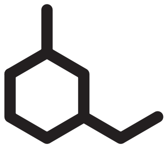
Cyclopropane



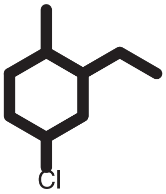
Cyclopentane

Naming substituted cycloalkanes is straightforward: We name them as *alkylcycloalkanes, halocycloalkanes, alkylcycloalkanols*, and so on. If only one substituent is present, it is not necessary to designate its position. When two substituents are present, we number the ring beginning with the substituent first in the alphabet and number in the direction that gives the next substituent the lower number possible. When three or more substituents are present, we begin at the substituent that leads to the lowest set of locants:

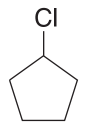
Isopropylcyclohexane



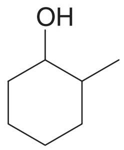
1-Ethyl-3-methylcyclohexane
(*not* 1-ethyl-5-methylcyclohexane)



4-Chloro-2-ethyl-1-methylcyclohexane
(*not* 1-chloro-3-ethyl-4-methylcyclohexane)

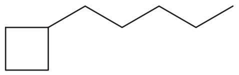


Chlorocyclopentane

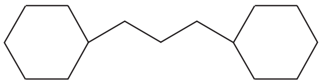


2-Methylcyclohexanol

When a single ring system is attached to a single chain with a greater number of carbon atoms, or when more than one ring system is attached to a single chain, then it is appropriate to name the compounds as *cycloalkylalkanes*. For example,



1-Cyclobutylpentane

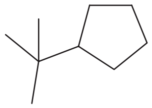


1,3-Dicyclohexylpropane

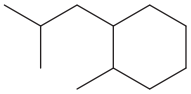
Review Problem 4.8

Give names for the following substituted alkanes:

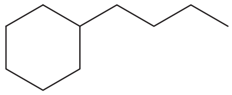
(a)



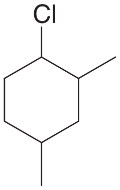
(b)



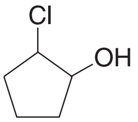
(c)



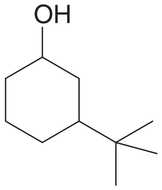
(d)



(e)



(f)

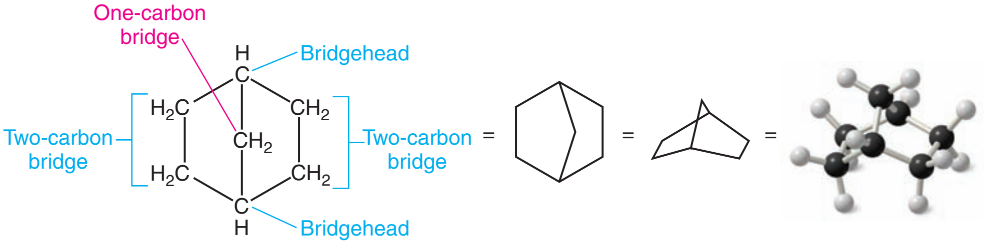


4.4B Bicyclic Compounds

We name compounds containing two fused or bridged rings as **bicycloalkanes**, and we use the name of the alkane corresponding to the total number of carbon atoms in the rings as the parent name. The following compound, for example, contains seven carbon atoms and is, therefore, a bicycloheptane. The carbon atoms common to both rings are called bridgeheads, and each bond, or each chain of atoms, connecting the bridgehead atoms is called a bridge:

Helpful Hint

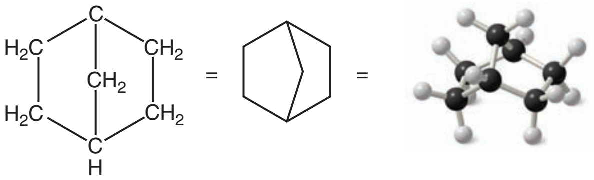
Explore the structures of these bicyclic compounds by building handheld molecular models.



A bicycloheptane

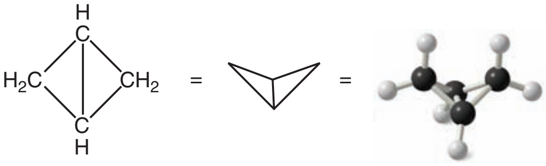
Then we interpose an expression in brackets within the name that denotes the number of carbon atoms in each bridge (in order of decreasing length). Fused rings have zero carbons in the bridge. For example,

Bridged



Bicyclo[2.2.1]heptane (also called *norbornane*)

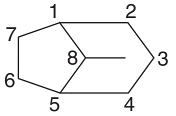
Fused



Bicyclo[1.1.0]butane

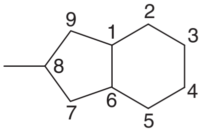
If substituents are present, we number the bridged ring system beginning at one bridgehead, proceeding first along the longest bridge to the other bridgehead, then along the next longest bridge back to the first bridgehead. The shortest bridge is numbered last:

Bridged



8-Methylbicyclo[3.2.1]octane

Fused

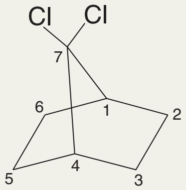


8-Methylbicyclo[4.3.0]nonane

Solved Problem 4.3

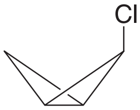
Write a structural formula for 7,7-dichlorobicyclo[2.2.1]heptane.

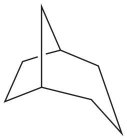
STRATEGY AND ANSWER First we write a bicyclo[2.2.1]heptane ring and number it. Then we add the substituents (two chlorine atoms) to the proper carbon.

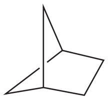


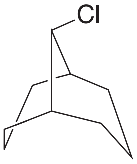
Review Problem 4.9

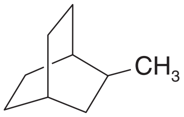
Give names for each of the following bicyclic alkanes:

(a)


(b)


(c)


(d)


(e)


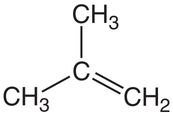
(f) Write the structure of a bicyclic compound that is a constitutional isomer of bicyclo[2.2.0]hexane and give its name.

4.5 *Nomenclature of Alkenes and Cycloalkenes*

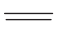
Many older names for alkenes are still in common use. Propene is often called propylene, and 2-methylpropene frequently bears the name isobutylene:

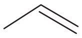
$CH_2{=}CH_2$

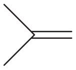
$CH_3CH{=}CH_2$

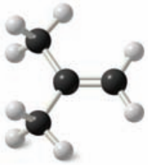


$$(CH_3)_2C{=}CH_2$$









IUPAC:	**Ethene**	**Propene**	**2-Methylpropene**
Common:	***Ethylene***	***Propylene***	***Isobutylene***

The IUPAC rules for naming alkenes are similar in many respects to those for naming alkanes:

1. **Determine the parent name by selecting the longest chain that contains the double bond and change the ending of the name of the alkane of identical length from -ane to -ene**. Thus, if the longest chain contains five carbon atoms, the parent name for the alkene is *pentene*; if it contains six carbon atoms, the parent name is *hexene*, and so on.
2. **Number the chain so as to include both carbon atoms of the double bond, and begin numbering at the end of the chain nearer the double bond. Designate the location of the double bond by using the number of the first atom of the double bond as a prefix. The locant for the alkene suffix may precede the parent name or be placed immediately before the suffix**. We will show examples of both styles:

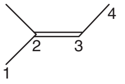
$\overset{1}{C}H_2=\overset{2}{C}H\overset{3}{C}H_2\overset{4}{C}H_3$

1-Butene
(*not* 3-butene)

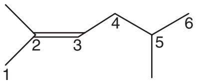
$CH_3CH=CHCH_2CH_2CH_3$

2-Hexene
(*not* 4-hexene)

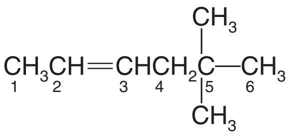
3. **Indicate the locations of the substituent groups by the numbers of the carbon atoms to which they are attached**:



2-Methyl-2-butene
or 2-methylbut-2-ene



2,5-Dimethyl-2-hexene
or 2,5-dimethylhex-2-ene



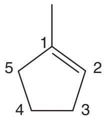
$$\underset{1}{C}H_3\underset{2}{C}H=\underset{3}{C}H\underset{4}{C}H_2\overset{\displaystyle CH_3}{\underset{\displaystyle CH_3}{\overset{|}{\underset{|\,5}{C}}}}-\underset{6}{C}H_3$$

5,5-Dimethyl-2-hexene
or 5,5-dimethylhex-2-ene

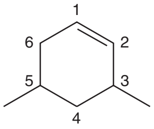
$\overset{4}{C}H_3\overset{3}{C}H=\overset{2}{C}H\overset{1}{C}H_2Cl$

1-Chloro-2-butene
or 1-chlorobut-2-ene

4. **Number substituted cycloalkenes in the way that gives the carbon atoms of the double bond the 1 and 2 positions and that also gives the substituent groups the lower numbers at the first point of difference**. With substituted cycloalkenes it is not necessary to specify the position of the double bond since it will always begin with C1 and C2. The two examples shown here illustrate the application of these rules:

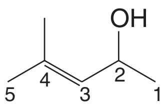


1-Methylcyclopentene
(*not* 2-methylcyclopentene)

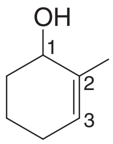


3,5-Dimethylcyclohexene
(*not* 4,6-dimethylcyclohexene)

5. Name compounds containing a double bond and an alcohol group as alkenols (or cycloalkenols) and give the alcohol carbon the lower number:



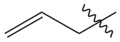
4-Methyl-3-penten-2-ol
or 4-methylpent-3-en-2-ol



2-Methyl-2-cyclohexen-1-ol
or 2-methylcyclohex-2-en-1-ol

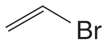
6. Two frequently encountered alkenyl groups are the **vinyl group** and the **allyl group**:

The vinyl group

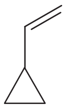


The allyl group

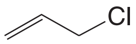
Using substitutive nomenclature, the vinyl and allyl groups are called *ethenyl* and *prop-2-en-1-yl*, respectively. The following examples illustrate how these names are employed:



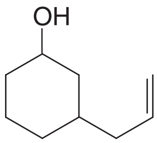
Bromoethene
or
vinyl bromide
(common)



Ethenylcyclopropane
or
vinylcyclopropane

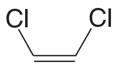


3-Chloropropene
or
allyl chloride
(common)

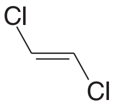


3-(Prop-2-en-1-yl)cyclohexan-1-ol
or
3-allylcyclohexanol

7. If two identical or substantial groups are on the same side of the double bond, the compound can be designated *cis*; if they are on opposite sides it can be designated *trans*:



***cis*-1,2-Dichloroethene**

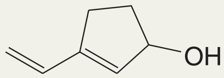


***trans*-1,2-Dichloroethene**

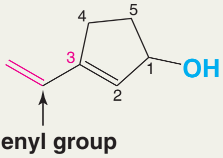
(In Section 7.2 we shall see another method for designating the geometry of the double bond.)

Solved Problem 4.4

Give an IUPAC name for the following molecule.



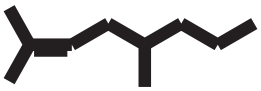
STRATEGY AND ANSWER We number the ring as shown below starting with the hydroxyl group so as to give the double bond the lower possible number. We include in the name the substituent (an ethenyl group) and the double bond (*-ene-*), and the hydroxyl group (*-ol*) with numbers for their respective positions. Hence the IUPAC name is 3-ethenyl-2-cyclopenten-1-ol.



Review Problem 4.10

Give IUPAC names for the following alkenes:

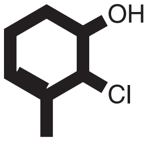
(a)


(b)


(c)


(d)


(e)


(f)


Review Problem 4.11

Write bond-line formulas for the following:

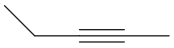
(a) *cis*-3-Octene
(b) *trans*-2-Hexene
(c) 2,4-Dimethyl-2-pentene
(d) *trans*-1-Chlorobut-2-ene
(e) 4,5-Dibromo-1-pentene
(f) 1-Bromo-2-methyl-1-(prop-2-en-1-yl)cyclopentane
(g) 3,4-Dimethylcyclopentene
(h) Vinylcyclopentane
(i) 1,2-Dichlorocyclohexene
(j) *trans*-1,4-Dichloro-2-pentene

4.6 Nomenclature of Alkynes

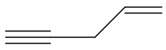
Alkynes are named in much the same way as alkenes. Unbranched alkynes, for example, are named by replacing the **-ane** of the name of the corresponding alkane with the ending **-yne.** The chain is numbered to give the carbon atoms of the triple bond the lower possible numbers. The lower number of the two carbon atoms of the triple bond is used to designate the location of the triple bond. The IUPAC names of three unbranched alkynes are shown here:

$H-C\equiv C-H$

Ethyne or acetylene*

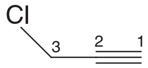


2-Pentyne

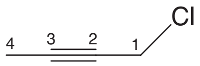


1-Penten-4-yne[†] **or pent-1-en-4-yne**

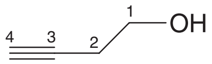
The locations of substituent groups of branched alkynes and substituted alkynes are also indicated with numbers. An —OH group has priority over the triple bond when numbering the chain of an alkynol:



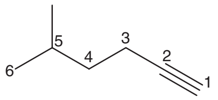
3-Chloropropyne



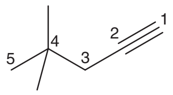
1-Chloro-2-butyne or 1-chlorobut-2-yne



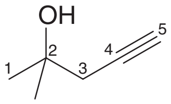
3-Butyn-1-ol or but-3-yn-1-ol



5-Methyl-1-hexyne or 5-methylhex-1-yne



4,4-Dimethyl-1-pentyne or 4,4-dimethylpent-1-yne



2-Methyl-4-pentyn-2-ol or 2-methylpent-4-yn-2-ol

Review Problem 4.12 Give the structures and IUPAC names for all the alkynes with the formula C_6H_{10}.

Monosubstituted acetylenes or 1-alkynes are called **terminal alkynes**, and the hydrogen attached to the carbon of the triple bond is called the **acetylenic hydrogen atom**:

$R-C\equiv C-H$

Acetylenic hydrogen

A terminal alkyne

When named as a substituent, the $HC\equiv C-$ group is called the ethynyl group.

The anion obtained when the acetylenic hydrogen is removed is known as an *alkynide ion* or an acetylide ion. As we shall see in Section 7.11, these ions are useful in synthesis:

$R-C\equiv C:^-$

or

$R-\equiv:^-$

An alkynide ion (an acetylide ion)

$CH_3C\equiv C:^-$

or

$-\equiv:^-$

The propynide ion

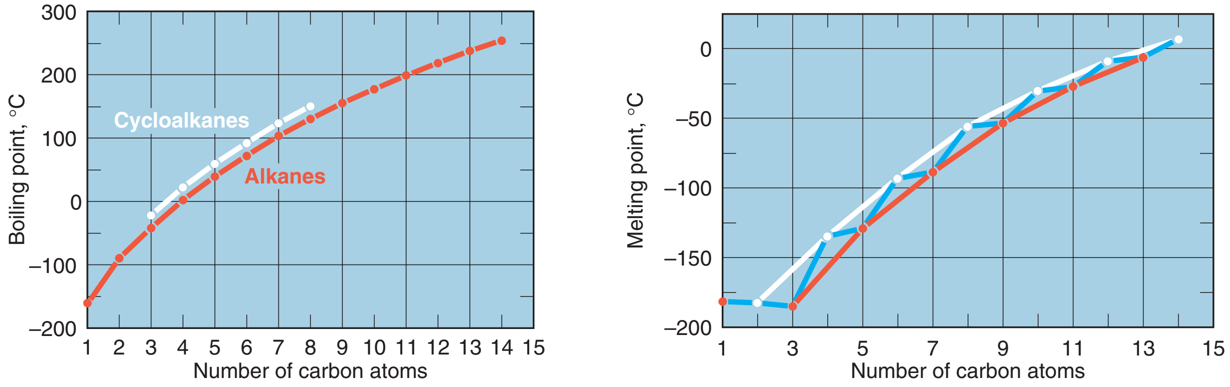
4.7 Physical Properties of Alkanes and Cycloalkanes

If we examine the unbranched alkanes in Table 4.3, we notice that each alkane differs from the preceding alkane by one $-CH_2-$ group. Butane, for example, is $CH_3(CH_2)_2CH_3$ and pentane is $CH_3(CH_2)_3CH_3$. A series of compounds like this, where each member differs

---

*The name acetylene is retained by the IUPAC system for the compound $HC\equiv CH$ and is used frequently.

†Where there is a choice, the double bond is given the lower number.



(a) (b)

Figure 4.3 *(a)* Boiling points of unbranched alkanes (in red) and cycloalkanes (in white). *(b)* Melting points of unbranched alkanes.

from the next member by a constant unit, is called a **homologous series**. Members of a homologous series are called **homologues**.

At room temperature (25°C) and 1 atm pressure the first four members of the homologous series of unbranched alkanes are gases (Fig. 4.3), the C_5—C_{17} unbranched alkanes (pentane to heptadecane) are liquids, and the unbranched alkanes with 18 and more carbon atoms are solids.

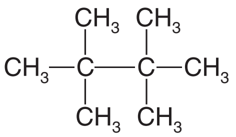
Boiling Points The boiling points of the unbranched alkanes show a regular increase with increasing molecular weight (Fig. 4.3*a*) in the homologous series of straight-chain alkanes. Branching of the alkane chain, however, lowers the boiling point. As examples, consider the C_6H_{14} isomers in Table 4.1. Hexane boils at 68.7°C and 2-methylpentane and 3-methylpentane, each having one branch, boil lower, at 60.3 and 63.3°C, respectively. 2,3-Dimethylbutane and 2,2-dimethylbutane, each with two branches, boil lower still, at 58 and 49.7°C, respectively.

Part of the explanation for these effects lies in the dispersion forces that we studied in Section 2.13B. With unbranched alkanes, as molecular weight increases, so too do molecular size and, even more importantly, molecular surface area. With increasing surface area, the dispersion forces between molecules increase; therefore, more energy (a higher temperature) is required to separate molecules from one another and produce boiling. Chain branching, on the other hand, makes a molecule more compact, reducing its surface area and with it the strength of the dispersion forces operating between it and adjacent molecules; this has the effect of lowering the boiling point.

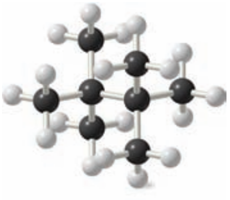
Melting Points The unbranched alkanes do not show the same smooth increase in melting points with increasing molecular weight (blue line in Fig. 4.3*b*) that they show in their boiling points. There is an alternation as one progresses from an unbranched alkane with an even number of carbon atoms to the next one with an odd number of carbon atoms. For example, propane (mp −188°C) melts lower than ethane (mp −183°C) and also lower than methane (mp −182°C). Butane (mp −138°C) melts 50°C higher than propane and only 8°C lower than pentane (mp −130°C). If, however, the even- and odd-numbered alkanes are plotted on *separate* curves (white and red lines in Fig. 4.3*b*), there *is* a smooth increase in melting point with increasing molecular weight.

X-Ray diffraction studies, which provide information about molecular structure, have revealed the reason for this apparent anomaly. Alkane chains with an even number of carbon atoms pack more closely in the crystalline state. As a result, attractive forces between individual chains are greater and melting points are higher.

The effect of chain branching on the melting points of alkanes is more difficult to predict. Generally, however, branching that produces highly symmetrical structures results in abnormally high melting points. The compound 2,2,3,3-tetramethylbutane, for example, melts at 100.7°C. Its boiling point is only six degrees higher, 106.3°C:



2,2,3,3-Tetramethylbutane



Cycloalkanes also have much higher melting points than their open-chain counterparts (Table 4.4).

TABLE 4.4 Physical Constants of Cycloalkanes

Number of Carbon Atoms	Name	bp (°C) (1 atm)	mp (°C)	Density, d^{20}(g mL^{-1})	Refractive Index (n_D^{20})
3	Cyclopropane	−33	−126.6	—	—
4	Cyclobutane	13	−90	—	1.4260
5	Cyclopentane	49	−94	0.751	1.4064
6	Cyclohexane	81	6.5	0.779	1.4266
7	Cycloheptane	118.5	−12	0.811	1.4449
8	Cyclooctane	149	13.5	0.834	—

Density As a class, the alkanes and cycloalkanes are the least dense of all groups of organic compounds. All alkanes and cycloalkanes have densities considerably less than 1.00 g mL^{-1} (the density of water at 4°C). As a result, petroleum (a mixture of hydrocarbons rich in alkanes) floats on water.

Solubility Alkanes and cycloalkanes are almost totally insoluble in water because of their very low polarity and their inability to form hydrogen bonds. Liquid alkanes and cycloalkanes are soluble in one another, and they generally dissolve in solvents of low polarity. Good solvents for them are benzene, carbon tetrachloride, chloroform, and other hydrocarbons.

THE CHEMISTRY OF . . .

Pheromones: Communication by Means of Chemicals

Many animals communicate with other members of their species using a language based not on sounds or even visual signals but on the odors of chemicals called **pheromones** that these animals release. For insects, this appears to be the chief method of communication. Although pheromones are secreted by insects in extremely small amounts, they can cause profound and varied biological effects. Some insects use pheromones in courtship as sex attractants. Others use pheromones as warning substances, and still others secrete chemicals called "aggregation compounds" to cause members of their species to congregate. Often these pheromones are relatively simple compounds, and some are hydrocarbons. For example, a species of cockroach uses undecane as an aggregation pheromone:

$CH_3(CH_2)_9CH_3$

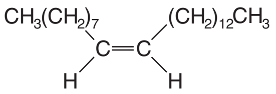
Undecane
(cockroach aggregation pheromone)

$(CH_3)_2CH(CH_2)_{14}CH_3$

2-Methylheptadecane
(sex attractant of female tiger moth)

When a female tiger moth wants to mate, she secretes 2-methylheptadecane, a perfume that the male tiger moth apparently finds irresistible.

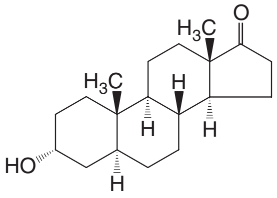
The sex attractant of the common housefly *(Musca domestica)* is a 23-carbon alkene with a cis double bond between atoms 9 and 10 called muscalure:



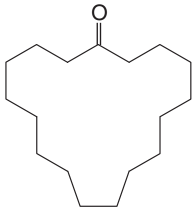
Muscalure
(sex attractant of common housefly)

Many insect sex attractants have been synthesized and are used to lure insects into traps as a means of insect control, a much more environmentally sensitive method than the use of insecticides.

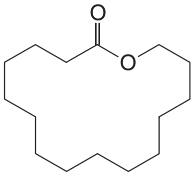
Research suggests there are roles for pheromones in the lives of humans as well. For example, studies have shown that the phenomenon of menstrual synchronization among women who live or work with each other is likely caused by pheromones. Olfactory sensitivity to musk, which includes steroids such as androsterone, large cyclic ketones, and lactones (cyclic esters), also varies cyclically in women, differs between the sexes, and may influence our behavior. Some of these compounds are used in perfumes, including civetone, a natural product isolated from glands of the civet cat, and pentalide, a synthetic musk.



Androsterone



Civetone



Pentalide

4.8 *Sigma Bonds and Bond Rotation*

Two groups bonded by only a single bond can undergo rotation about that bond with respect to each other.

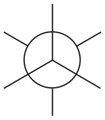
- The temporary molecular shapes that result from such a rotation are called **conformations** of the molecule.
- Each possible structure is called a **conformer**.
- An analysis of the energy changes that occur as a molecule undergoes rotations about single bonds is called a **conformational analysis**.

4.8A Newman Projections and How to Draw Them

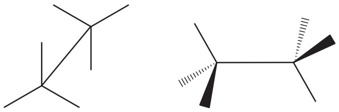
When we do conformational analysis, we will find that certain types of structural formulas are especially convenient to use. One of these types is called a **Newman projection formula** and another type is a **sawhorse formula**. Sawhorse formulas are much like dash–wedge three-dimensional formulas we have used so far. In conformational analyses, we will make substantial use of Newman projections.

Helpful Hint

Learn to draw Newman projections and sawhorse formulas. Build handheld molecular models and compare them with your drawings.

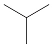
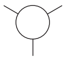


Newman projection formula



Sawhorse formula

To write a Newman projection formula:

- We imagine ourselves taking a view from one atom (usually a carbon) directly along a selected bond axis to the next atom (also usually a carbon atom).
- The front carbon and its other bonds are represented as .
- The back carbon and its bonds are represented as .

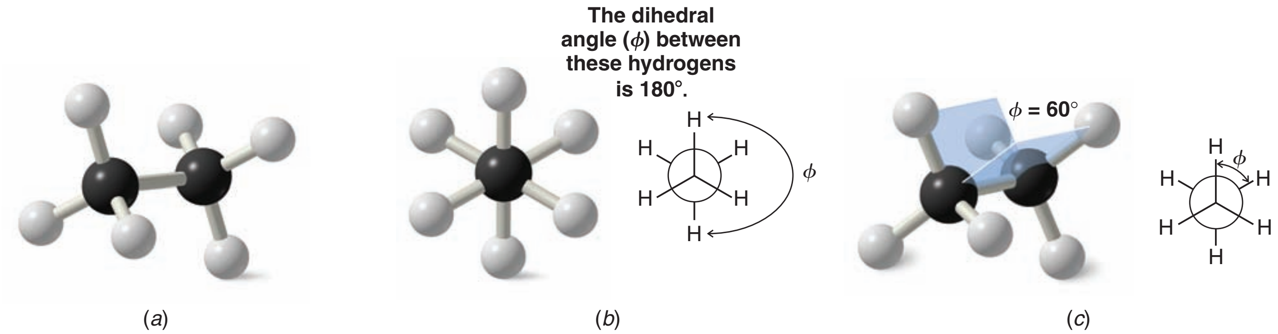


Figure 4.4 *(a)* The staggered conformation of ethane. *(b)* The Newman projection formula for the staggered conformation. *(c)* The dihedral angle between these hydrogen atoms is 60°.

In Figs. 4.4*a*,*b* we show ball-and-stick models and a Newman projection formula for the staggered conformation of ethane. The **staggered conformation** of a molecule is that conformation where the **dihedral angle** between the bonds at each of the carbon–carbon bonds is 180° and where atoms or groups bonded to carbons at each end of a carbon–carbon bond are as far apart as possible. The 180° dihedral angle in the staggered conformation of ethane is indicated in Fig. 4.4*b*.

The eclipsed conformation of ethane is shown in Fig. 4.5 using ball-and-stick models and a Newman projection. In an **eclipsed conformation** the atoms bonded to carbons at each end of a carbon–carbon bond are directly opposed to one another. The dihedral angle between them is 0°.

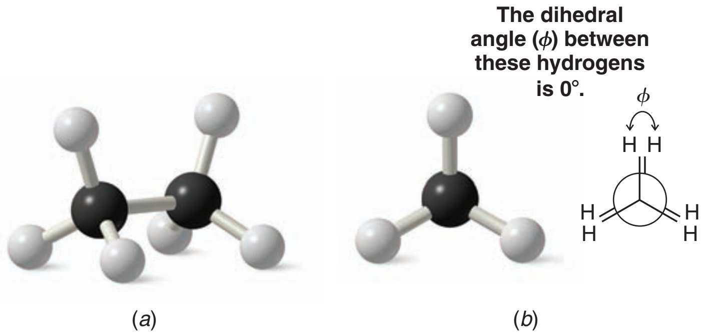


Figure 4.5 *(a)* The eclipsed conformation of ethane. *(b)* The Newman projection formula for the eclipsed conformation.

4.8B How to Do a Conformational Analysis

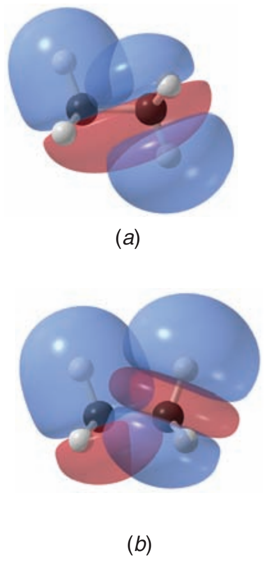
Now let us consider a conformational analysis of ethane. Clearly, infinitesimally small changes in the dihedral angle between C—H bonds at each end of ethane could lead to an infinite number of conformations, including, of course, the staggered and eclipsed conformations. These different conformations are not all of equal stability, however, and it is known that the staggered conformation of ethane is the most stable conformation (i.e., it is the conformation of lowest potential energy). The fundamental reason for this has recently come to light.

Quantum mechanical calculations by L. Goodman and V. T. Pophristic (Rutgers University) have shown that the greater stability of the staggered conformation in ethane over the eclipsed conformation is mainly due to favorable overlap between sigma (σ) bonding orbitals from the C—H bonds at one carbon and unfilled antibonding sigma (σ^*) orbitals at the adjacent carbon. In ethane's staggered conformation, electrons from a given bonding C—H orbital on one carbon can be shared with an unfilled σ^* orbital at the adjacent carbon. This phenomenon of electron delocalization (via orbital overlap) from a filled bonding orbital to an adjacent unfilled orbital is called **hyperconjugation** (and we shall see in later chapters that it is a general stabilizing effect). Figure 4.6*a* shows the favorable overlap of σ and σ^* in ethane by color coding of the orbital phases.

If we now consider the eclipsed conformation of ethane (Fig. 4.6*b*), where the C—H bonds at each carbon are directly opposed to each other, we see that the bonding σ C—H orbital at one carbon does not overlap to as great an extent with the adjacent antibonding orbital as in the staggered conformation. The possibility for hyperconjugation is diminished, and therefore the potential energy of this conformation is higher.

The σ–σ^* interactions in ethane are present in more complicated molecules as well. However, where atoms and groups larger than hydrogen are involved in a conformational analysis, such as our example in Section 4.9, it is likely that repulsion of the electron clouds involved in the bonding of those groups increases in importance as the cause of the staggered conformation being most stable.

The energy difference between the conformations of ethane can be represented graphically in a **potential energy diagram**, as shown in Figure 4.7. In ethane the energy difference between the staggered and eclipsed conformations is about 12 kJ mol^{-1}. This small barrier to rotation is called the **torsional barrier** of the single bond. Because of this barrier, some molecules will wag back and forth with their atoms in staggered or nearly staggered conformations, while others with slightly more energy will rotate through an eclipsed conformation to another staggered conformation. At any given moment, unless the temperature is extremely low (−250°C), most ethane molecules will have enough energy to undergo bond rotation from one conformation to another.


(*a*)

(*b*)

Figure 4.6 The bonding C—H orbital is the orbital where one lobe of a single phase (represented by one color) envelops the C—H atoms. The adjacent unfilled antibonding orbital is the orbital where the phase changes between the carbon and its hydrogen atom. The staggered conformation of ethane *(a)* has greater overlap between the bonding C—H orbital and the adjacent antibonding orbital than in the eclipsed conformation *(b)*. The orbital overlap shown in *(a)* leads to the lower potential energy of the staggered conformation of ethane.

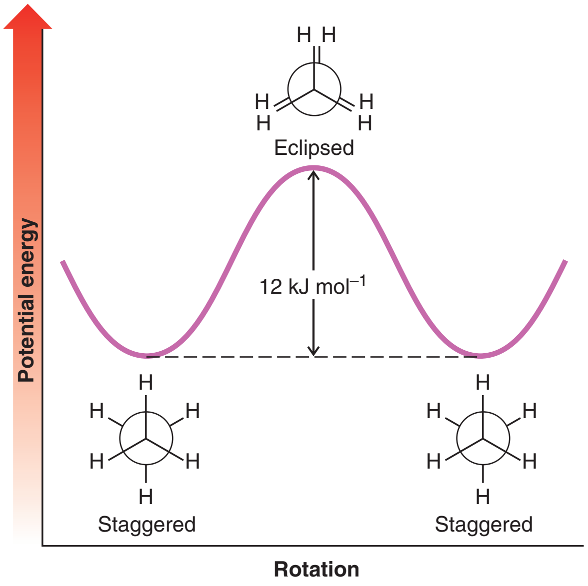


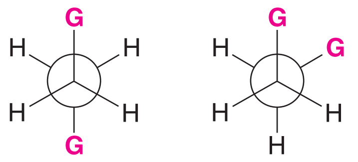
Figure 4.7 Potential energy changes that accompany rotation of groups about the carbon–carbon bond of ethane.

What does all this mean about ethane? We can answer this question in two different ways. If we consider a single molecule of ethane, we can say, for example, that it will spend most of its time in the lowest energy, staggered conformation, or in a conformation very close to being staggered. Many times every second, however, it will acquire enough energy through collisions with other molecules to surmount the torsional barrier and it will rotate through an eclipsed conformation. If we speak in terms of a large number of ethane molecules (a more realistic situation), we can say that at any given moment most of the molecules will be in staggered or nearly staggered conformations.

If we consider more highly substituted ethanes such as GCH_2CH_2G (where G is a group or atom other than hydrogen), the barriers to rotation are somewhat larger, but they are still far too small to allow isolation of the different staggered conformations. The factors involved in this rotational barrier are together called **torsional strain** and include the orbital considerations discussed above as well as repulsive interactions called **steric**

The idea that certain conformations of molecules are favored originates from the work of J.H. van't Hoff. He was also winner of the first Nobel Prize in Chemistry (1901) for his work in chemical kinetics.

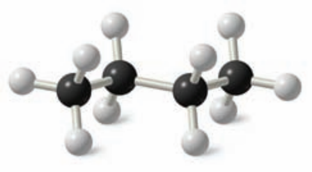
hindrance between electron clouds of bonded groups. In the next section we consider a conformational analysis of butane, where groups larger than hydrogen are involved in the analysis.



Conformers like these cannot be isolated except at extremely low temperatures.

4.9 *Conformational Analysis of Butane*

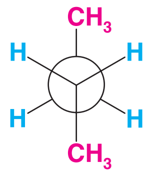
If we consider rotation about the C2—C3 bond of butane, we find that there are six important conformers, shown as **I–VI** below:



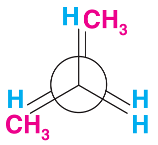
Butane

Helpful Hint

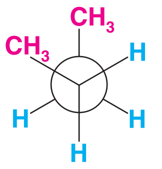
You should build a molecular model of butane and examine its various conformations as we discuss their relative potential energies.



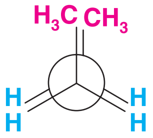
I
An anti conformation



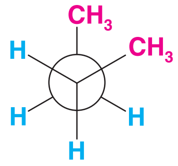
II
An eclipsed conformation



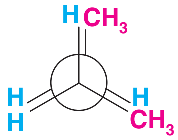
III
A gauche conformation



IV
An eclipsed conformation



V
A gauche conformation



VI
An eclipsed conformation

The **anti conformation** (**I**) does not have torsional strain from steric hindrance because the groups are staggered and the methyl groups are far apart. The anti conformation is the most stable. The methyl groups in the **gauche conformations** **III** and **V** are close enough to each other that the dispersion forces between them are *repulsive*; the electron clouds of the two groups are so close that they repel each other. This repulsion causes the gauche conformations to have approximately 3.8 kJ mol^{-1} more energy than the anti conformation.

The eclipsed conformations (**II, IV**, and **VI**) represent energy maxima in the potential energy diagram (Fig. 4.8). Eclipsed conformations **II** and **VI** have repulsive dispersion forces arising from the eclipsed methyl groups and hydrogen atoms. Eclipsed conformation **IV** has the greatest energy of all because of the added large repulsive dispersion forces between the eclipsed methyl groups as compared to **II** and **VI**.

Although the barriers to rotation in a butane molecule are larger than those of an ethane molecule (Section 4.8), they are still far too small to permit isolation of the gauche and anti conformations at normal temperatures. Only at extremely low temperatures would the molecules have insufficient energies to surmount these barriers.

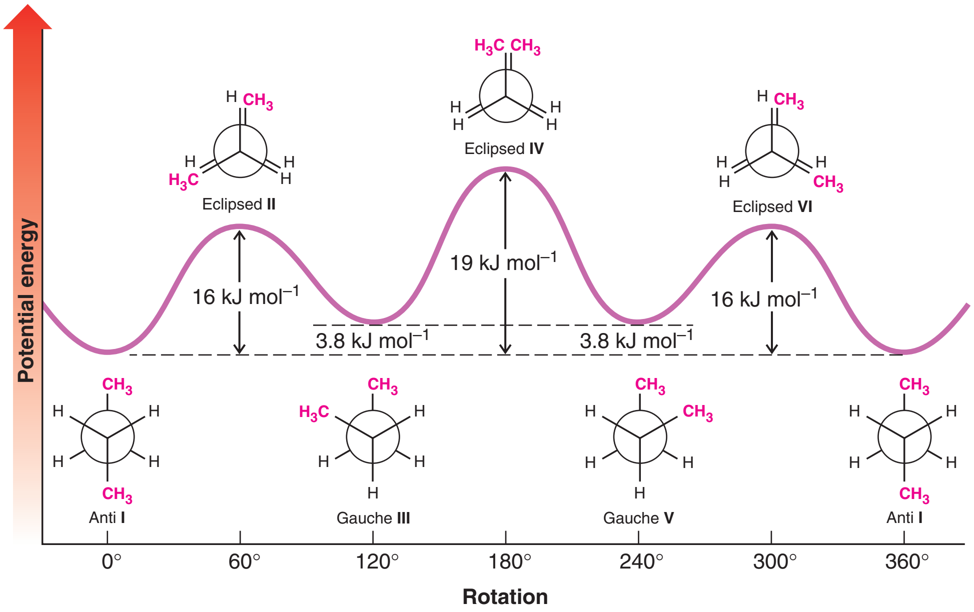


Figure 4.8 Energy changes that arise from rotation about the C2—C3 bond of butane.

We saw earlier (Section 2.16C) that dispersion forces can be *attractive*. Here, however, we find that they can also be *repulsive*, leading to steric hindrance. Whether dispersion interactions lead to attraction or to repulsion depends on the distance that separates the two groups. As two nonpolar groups are brought closer and closer together, the first effect is one in which a momentarily unsymmetrical distribution of electrons in one group induces an opposite polarity in the other. The opposite charges induced in those portions of the two groups that are in closest proximity lead to attraction between them. This attraction increases to a maximum as the internuclear distance of the two groups decreases. The internuclear distance at which the attractive force is at a maximum is equal to the sum of what are called the *van der Waals radii* of the two groups. The van der Waals radius of a group is, in effect, a measure of its size. If the two groups are brought still closer—closer than the sum of their van der Waals radii—the interaction between them becomes repulsive. Their electron clouds begin to penetrate each other, and strong electron–electron interactions begin to occur.

4.9A Stereoisomers and Conformational Stereoisomers

Gauche conformers **III** and **V** of butane are examples of stereoisomers.

- **Stereoisomers** have the same molecular formula and connectivity but different arrangements of atoms in three-dimensional space.
- **Conformational stereoisomers** are related to one another by bond rotations.

Conformational analysis is but one of the ways in which we will consider the three-dimensional shapes and stereochemistry of molecules. We shall see that there are other types of stereoisomers that cannot be interconverted simply by rotations about single bonds. Among these are cis–trans cycloalkane isomers (Section 4.13) and others that we shall consider in Chapter 5.

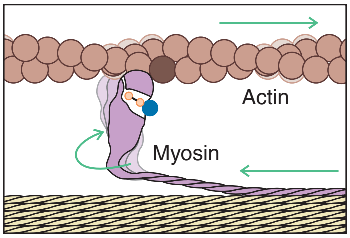
Review Problem 4.13

Sketch a curve similar to that in Fig. 4.8 showing in general terms the energy changes that arise from rotation about the C2—C3 bond of 2-methylbutane. You need not concern yourself with the actual numerical values of the energy changes, but you should label all maxima and minima with the appropriate conformations.

THE CHEMISTRY OF . . .

Muscle Action

Muscle proteins are essentially very long linear molecules (folded into a compact shape) whose atoms are connected by single bonds in a chainlike fashion. Relatively free rotation is possible about atoms joined by single bonds, as we have seen. In muscles, the cumulative effect of rotations about many single bonds is to move the tail of each myosin molecule 60 Å along the adjacent protein (called actin) in a step called the "power stroke." This process occurs over and over again as part of a ratcheting mechanism between many myosin and actin molecules for each muscle movement.



Power stroke in muscle

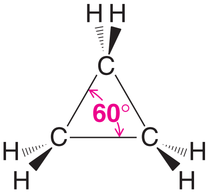
4.10 *The Relative Stabilities of Cycloalkanes: Ring Strain*

Cycloalkanes do not all have the same relative stability. Experiments have shown that cyclohexane is the most stable cycloalkane and that, in comparison, cyclopropane and cyclobutane are much less stable. This difference in relative stability is due to **ring strain**, which comprises **angle strain** and **torsional strain**.

- **Angle strain** is the result of deviation from ideal bond angles caused by inherent structural constraints (such as ring size).
- **Torsional strain** is the result of dispersion forces that cannot be relieved due to restricted conformational mobility.

4.10A Cyclopropane

The carbon atoms of alkanes are sp^3 hybridized. The normal tetrahedral bond angle of an sp^3-hybridized atom is 109.5°. In cyclopropane (a molecule with the shape of a regular triangle), the internal angles must be 60° and therefore they must depart from this ideal value by a very large amount—by 49.5°:



Angle strain exists in a cyclopropane ring because the sp^3 orbitals of the carbon atoms cannot overlap as effectively (Fig. 4.9*a*) as they do in alkanes (where perfect end-on overlap is possible). The carbon–carbon bonds of cyclopropane are often described as being "bent." Orbital overlap is less effective. (The orbitals used for these bonds are not purely sp^3; they contain more p character.) The carbon–carbon bonds of cyclopropane are weaker, and as a result the molecule has greater potential energy.

While angle strain accounts for most of the ring strain in cyclopropane, it does not account for it all. Because the ring is (of necessity) planar, the C—H bonds of the ring are all *eclipsed* (Figs. 4.9*b*,*c*), and the molecule has torsional strain from repulsive dispersion forces as well.

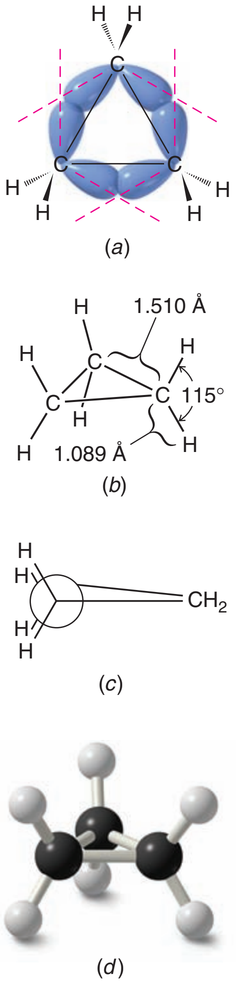
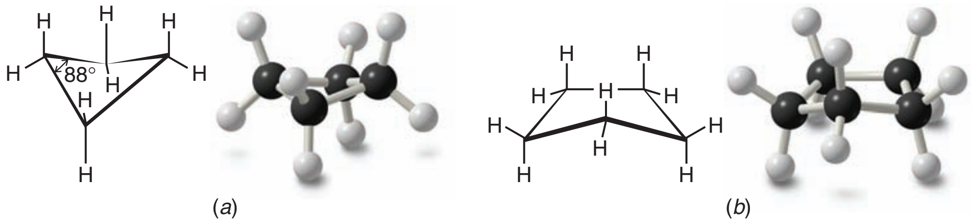


Figure 4.9 *(a)* Orbital overlap in the carbon–carbon bonds of cyclopropane cannot occur perfectly end-on. This leads to weaker "bent" bonds and to angle strain. *(b)* Bond distances and angles in cyclopropane. *(c)* A Newman projection formula as viewed along one carbon–carbon bond shows the eclipsed hydrogens. (Viewing along either of the other two bonds would show the same picture.) *(d)* Ball-and-stick model of cyclopropane.



(*a*) (*b*)

Figure 4.10 *(a)* The "folded" or "bent" conformation of cyclobutane. *(b)* The "bent" or "envelope" form of cyclopentane. In this structure the front carbon atom is bent upward. In actuality, the molecule is flexible and shifts conformations constantly.

4.10B Cyclobutane

Cyclobutane also has considerable angle strain. The internal angles are 88°—a departure of more than 21° from the normal tetrahedral bond angle. The cyclobutane ring is not planar but is slightly "folded" (Fig. 4.10*a*). If the cyclobutane ring were planar, the angle strain would be somewhat less (the internal angles would be 90° instead of 88°), but torsional strain would be considerably larger because all eight C—H bonds would be eclipsed. By folding or bending slightly the cyclobutane ring relieves more of its torsional strain than it gains in the slight increase in its angle strain.

4.10C Cyclopentane

The internal angles of a regular pentagon are 108°, a value very close to the normal tetrahedral bond angles of 109.5°. Therefore, if cyclopentane molecules were planar, they would have very little angle strain. Planarity, however, would introduce considerable torsional strain because all 10 C—H bonds would be eclipsed. Consequently, like cyclobutane, cyclopentane assumes a slightly bent conformation in which one or two of the atoms of the ring are out of the plane of the others (Fig. 4.10*b*). This relieves some of the torsional strain. Slight twisting of carbon–carbon bonds can occur with little change in energy and causes the out-of-plane atoms to move into plane and causes others to move out. Therefore, the molecule is flexible and shifts rapidly from one conformation to another. With little torsional strain and angle strain, cyclopentane is almost as stable as cyclohexane.

Helpful Hint

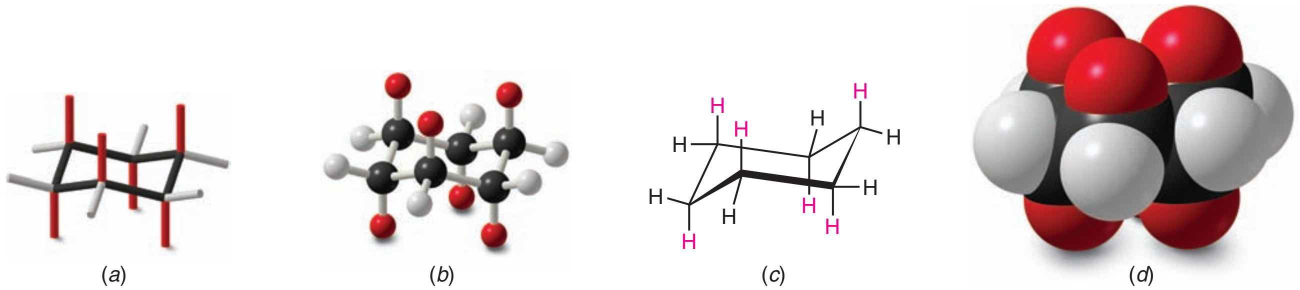
An understanding of this and subsequent discussions of conformational analysis can be aided immeasurably through the use of molecular models. We suggest you "follow along" with models as you read Sections 4.11–4.13.

4.11 *Conformations of Cyclohexane: The Chair and the Boat*

Cyclohexane is more stable than the other cycloalkanes we have discussed, and it has several conformations that are important for us to consider.

- The most stable conformation of cyclohexane is the **chair conformation**.
- There is no angle or torsional strain in the chair form of cyclohexane.

In a chair conformation (Fig. 4.11), all of the carbon–carbon bond angles are 109.5°, and are thereby free of angle strain. The chair conformation is free of torsional strain,



(*a*) (*b*) (*c*) (*d*)

Figure 4.11 Representations of the chair conformation of cyclohexane: *(a)* tube format; *(b)* ball-and-stick format; *(c)* line drawing; *(d)* space-filling model of cyclohexane. Notice that there are two types of hydrogen substituents—those that project obviously up or down (shown in red) and those that lie around the perimeter of the ring in more subtle up or down orientations (shown in black or gray). We shall discuss this further in Section 4.12.

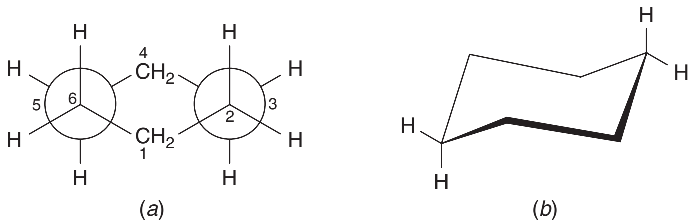


Figure 4.12 *(a)* A Newman projection of the chair conformation of cyclohexane. (Comparisons with an actual molecular model will make this formulation clearer and will show that similar staggered arrangements are seen when other carbon–carbon bonds are chosen for sighting.) *(b)* Illustration of large separation between hydrogen atoms at opposite corners of the ring (designated C1 and C4) when the ring is in the chair conformation.

as well. When viewed along any carbon–carbon bond (viewing the structure from an end, Fig. 4.12), the bonds are seen to be perfectly staggered. Moreover, the hydrogen atoms at opposite corners of the cyclohexane ring are maximally separated.

- By partial rotations about the carbon–carbon single bonds of the ring, the chair conformation can assume another shape called the **boat conformation** (Fig. 4.13).
- The boat conformation has no angle strain, but it does have torsional strain.

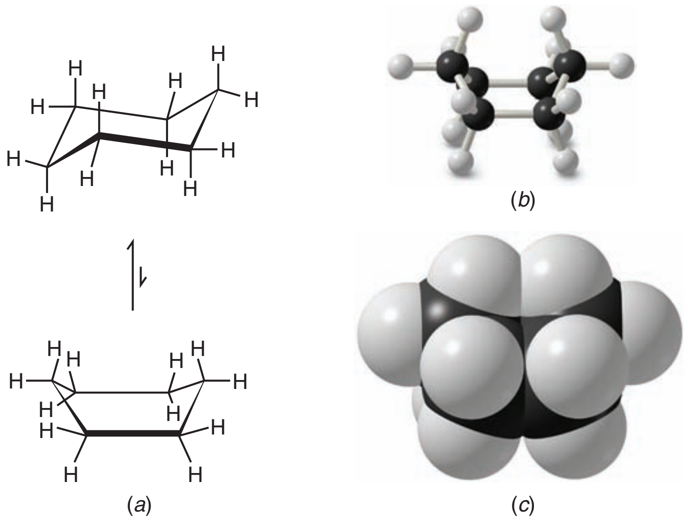


Figure 4.13 *(a)* The boat conformation of cyclohexane is formed by "flipping" one end of the chair form up (or down). This flip requires only rotations about carbon–carbon single bonds. *(b)* Ball-and-stick model of the boat conformation. *(c)* A space-filling model of the boat conformation.

Helpful Hint

You will best appreciate the differences between the chair and boat forms of cyclohexane by building and manipulating molecular models of each.

When a model of the boat conformation is viewed down carbon–carbon bond axes along either side (Fig. 4.14*a*), the C—H bonds at those carbon atoms are found to be eclipsed, causing torsional strain. Additionally, two of the hydrogen atoms on C1 and C4 are close enough to each other to cause van der Waals repulsion (Fig. 4.14*b*). This latter effect has

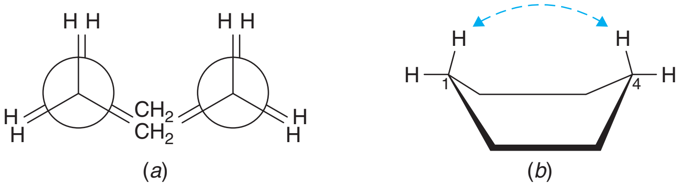


Figure 4.14 *(a)* Illustration of the eclipsed conformation of the boat conformation of cyclohexane. *(b)* Flagpole interaction of the C1 and C4 hydrogen atoms of the boat conformation. The C1–C4 flagpole interaction is also readily apparent in Fig. 4.13*c*.

been called the "flagpole" interaction of the boat conformation. Torsional strain and flagpole interactions cause the boat conformation to have considerably higher energy than the chair conformation.

Although it is more stable, the chair conformation is much more rigid than the boat conformation. The boat conformation is quite flexible. By flexing to a new form—the twist conformation (Fig. 4.15)—the boat conformation can relieve some of its torsional strain and, at the same time, reduce the flagpole interactions.

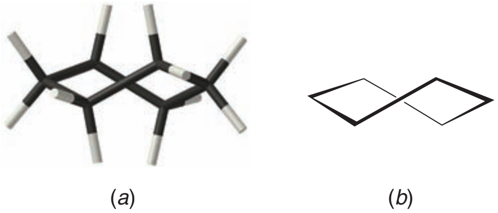


Figure 4.15 *(a)* Tube model and *(b)* line drawing of the twist conformation of cyclohexane.

- The twist boat conformation has a lower energy than the pure boat conformation, but is not as stable as the chair conformation.

The stability gained by flexing is insufficient, however, to cause the twist conformation of cyclohexane to be more stable than the chair conformation. The chair conformation is estimated to be lower in energy than the twist conformation by approximately 23 kJ mol^{-1}.

The energy barriers between the chair, boat, and twist conformations of cyclohexane are low enough (Fig. 4.16) to make separation of the conformers impossible at room temperature. At room temperature the thermal energies of the molecules are great enough to cause approximately 1 million interconversions to occur each second.

- *Because of the greater stability of the chair, more than 99% of the molecules are estimated to be in a chair conformation at any given moment.*

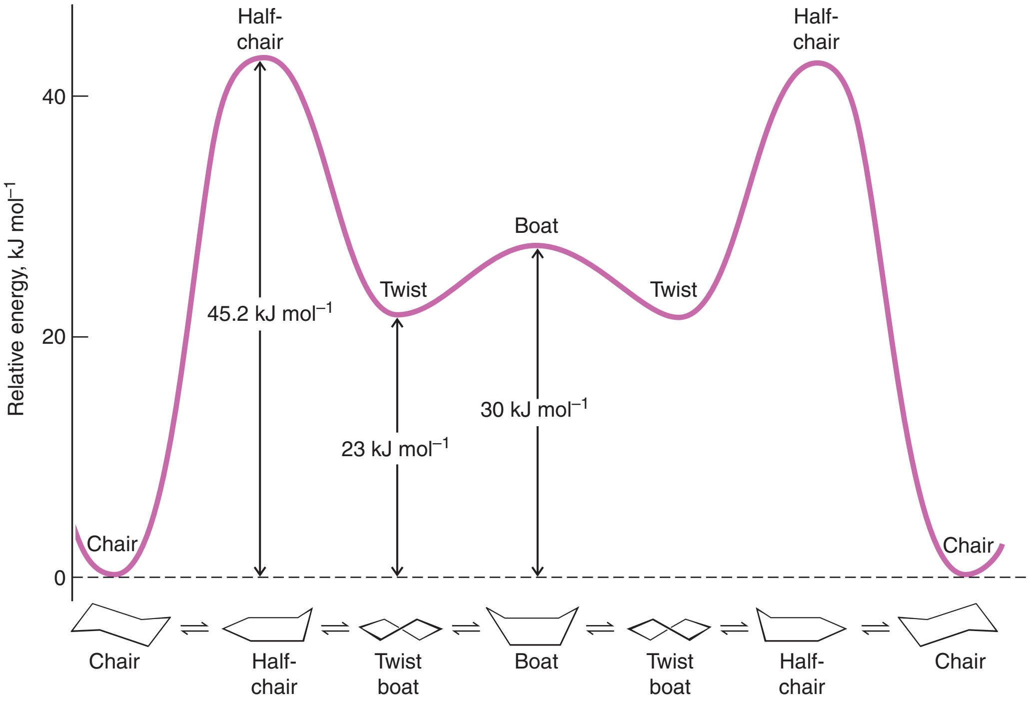


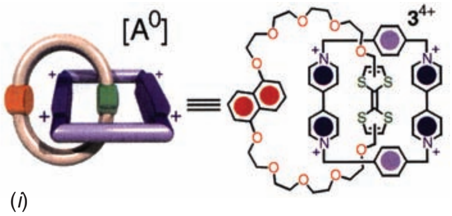
Figure 4.16 The relative energies of the various conformations of cyclohexane. The positions of maximum energy are conformations called half-chair conformations, in which the carbon atoms of one end of the ring have become coplanar.

THE CHEMISTRY OF . . .

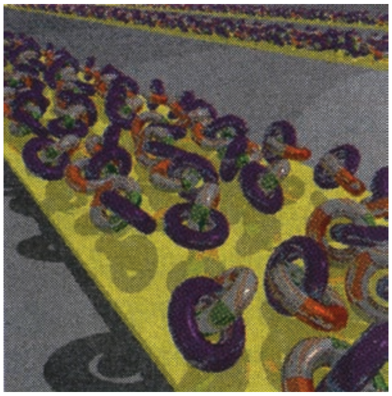
Nanoscale Motors and Molecular Switches

Molecular rings that interlock with one another and compounds that are linear molecules threaded through rings are proving to have fascinating potential for the creation of molecular switches and motors. Molecules consisting of interlocking rings, like a chain, are called **catenanes**. The first catenanes were synthesized in the 1960s and have come to include examples such as olympiadane, as mentioned in Section 4.11a. Further research by Stoddart (UCLA) and collaborators on interlocking molecules has led to examples such as the catenane molecular switch shown here in (i). In an application that could be useful in design of binary logic circuits, one ring of this molecule can be made to circumrotate in controlled fashion about the other, such that it switches between two defined states. As a demonstration of its potential for application in electronics fabrication, a monolayer of these molecules has been "tiled" on a surface (ii) and shown to have characteristics like a conventional magnetic memory bit.

Molecules where a linear molecule is threaded through a ring are called **rotaxanes**. One captivating example of a rotaxane system is the one shown here in (iii), under development by V. Balzani (University of Bologna) and collaborators. By conversion of light energy to mechanical energy at the molecular level, this rotaxane behaves like a "four-stroke" shuttle engine. In step (*a*) light excitation of an electron in the **P** group leads to transfer of the electron to the initially +2 $\mathbf{A_1}$ group, at which point $\mathbf{A_1}$ is reduced to the +1 state. Ring **R**, which was attracted to $\mathbf{A_1}$ when it was in the +2 state, now slides over to $\mathbf{A_2}$ in step (*b*), which remains +2. Back transfer of the electron from $\mathbf{A_1}$ to $\mathbf{P}^+$ in step (*c*) restores the +2 state of $\mathbf{A_1}$, causing ring **R** to return to its original location in step (*d*). Modifications envisioned for this system include attaching binding sites to **R** such that some other molecular species could be transported from one location to another as **R** slides along the linear molecule, or linking **R** by a springlike tether to one end of the "piston rod" such that additional potential and mechanical energy can be incorporated in the system.

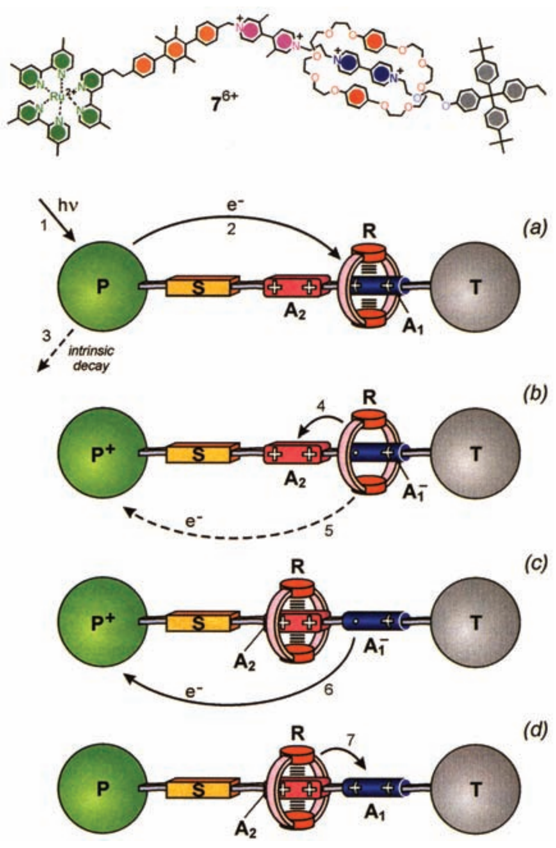


(*i*)

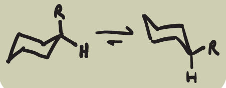


(*ii*)

(Figures reprinted with permission from Pease et al., *Accounts of Chemical Research*, *Vol. 34*, no. 6, pp. 433–444, 2001. Copyright 2001 American Chemical Society; and reprinted with permission from Ballardini et al., *Accounts of Chemical Research*, *Vol. 34*, no. 6, pp. 445–455, 2001. Copyright 2001 American Chemical Society.)



(*iii*)

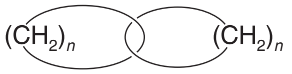


4.11A Conformations of Higher Cycloalkanes

Cycloheptane, cyclooctane, and cyclononane and other higher cycloalkanes also exist in nonplanar conformations. The small instabilities of these higher cycloalkanes appear to be caused primarily by torsional strain and repulsive dispersion forces between hydrogen atoms across rings, called *transannular strain*. The nonplanar conformations of these rings, however, are essentially free of angle strain.

X-Ray crystallographic studies of cyclodecane reveal that the most stable conformation has carbon–carbon–carbon bond angles of 117°. This indicates some angle strain. The wide bond angles apparently allow the molecules to expand and thereby minimize unfavorable repulsions between hydrogen atoms across the ring.

There is very little free space in the center of a cycloalkane unless the ring is quite large. Calculations indicate that cyclooctadecane, for example, is the smallest ring through which a —$CH_2CH_2CH_2$— chain can be threaded. Molecules have been synthesized, however, that have large rings threaded on chains and that have large rings that are interlocked like links in a chain. These latter molecules are called **catenanes**:



A catenane
($n \geq 18$)

In 1994 J. F. Stoddart and co-workers, then at the University of Birmingham (England), achieved a remarkable synthesis of a catenane containing a linear array of five interlocked rings. Because the rings are interlocked in the same way as those of the olympic symbol, they named the compound **olympiadane**.

Derek H. R. Barton (1918–1998) and Odd Hassel (1897–1981) shared the Nobel Prize in 1969 "for developing and applying the principles of conformation in chemistry." Their work led to fundamental understanding of not only the conformations of cyclohexane rings but also the structures of steroids (Section 23.4) and other compounds containing cyclohexane rings.

4.12 *Substituted Cyclohexanes: Axial and Equatorial Hydrogen Groups*

The six-membered ring is the most common ring found among nature's organic molecules. For this reason, we shall give it special attention. We have already seen that the chair conformation of cyclohexane is the most stable one and that it is the predominant conformation of the molecules in a sample of cyclohexane.

The chair conformation of a cyclohexane ring has two distinct orientations for the bonds that project from the ring. These positions are called axial and equatorial, as shown for cyclohexane in Fig. 4.17.

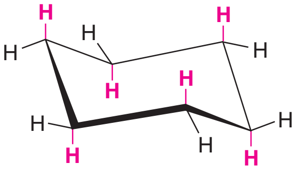
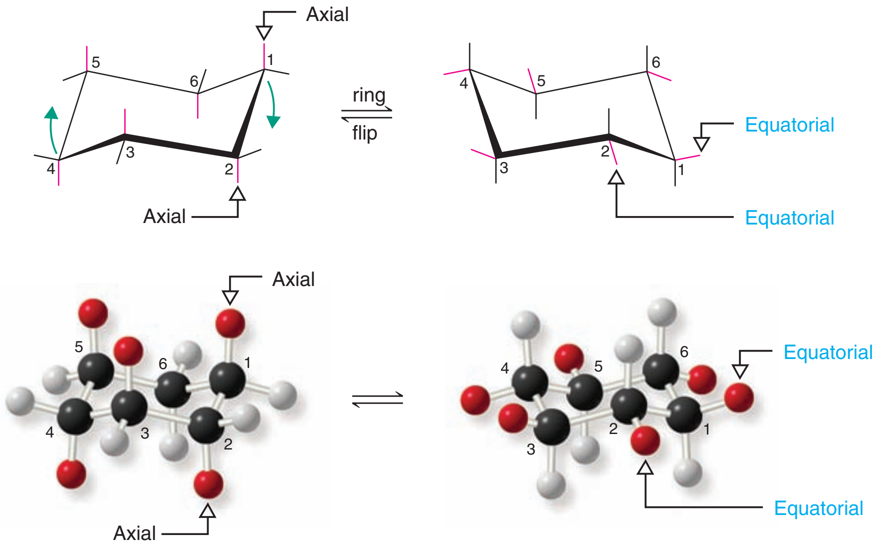


Figure 4.17 The chair conformation of cyclohexane. Axial hydrogen atoms are shown in red, equatorial hydrogens are shown in black.

- The **axial bonds** of cyclohexane are those that are perpendicular to the average plane of the ring. There are three axial bonds on each face of the cyclohexane ring, and their orientation (up or down) alternates from one carbon to the next.
- The **equatorial bonds** of cyclohexane are those that extend from the perimeter of the ring. The equatorial bonds alternate from slightly up to slightly down in their orientation from one carbon to the next.
- When a cyclohexane ring undergoes a chair–chair conformational change (a **ring flip**), all of the bonds that were axial become equatorial, and all bonds that were equatorial become axial.



4.12A How to Draw Chair Conformational Structures

A set of guidelines will help you draw chair conformational structures that are clear and that have unambiguous axial and equatorial bonds.

- Notice in Fig. 4.18*a* that sets of parallel lines define opposite sides of the chair. Notice, too, that equatorial bonds are parallel to ring bonds that are one bond away from them in either direction. When you draw chair conformational structures, try to make the corresponding bonds parallel in your drawings.
- When a chair formula is drawn as shown in Fig. 4.18, the axial bonds are all either up or down, in a vertical orientation (Fig. 4.18*b*). When a vertex of bonds in the ring points up, the axial bond at that position is also up, and the equatorial bond at the same carbon is angled slightly down. When a vertex of ring bonds is down, the axial bond at that position is also down, and the equatorial bond is angled slightly upward.

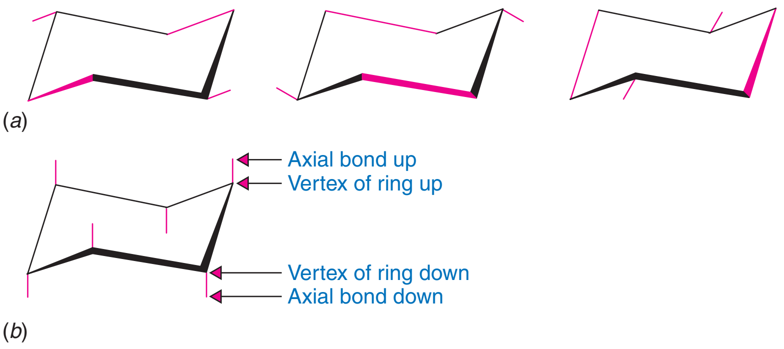


Figure 4.18 *(a)* Sets of parallel lines that constitute the ring and equatorial C—H bonds of the chair conformation. *(b)* The axial bonds are all vertical. When the vertex of the ring points up, the axial bond is up and vice versa.

Now, try to draw some chair conformational structures for yourself that include the axial and equatorial bonds. Then, compare your drawings with those here and with actual models. You will see that with a little practice your chair conformational structures can be perfect.

4.12B A Conformational Analysis of Methylcyclohexane

Now let us consider methylcyclohexane. Methylcyclohexane has two possible chair conformations (Fig. 4.19), and these are interconvertible through the bond rotations that constitute a ring flip. In one conformation (Fig. 4.19*a*) the methyl group (with yellow

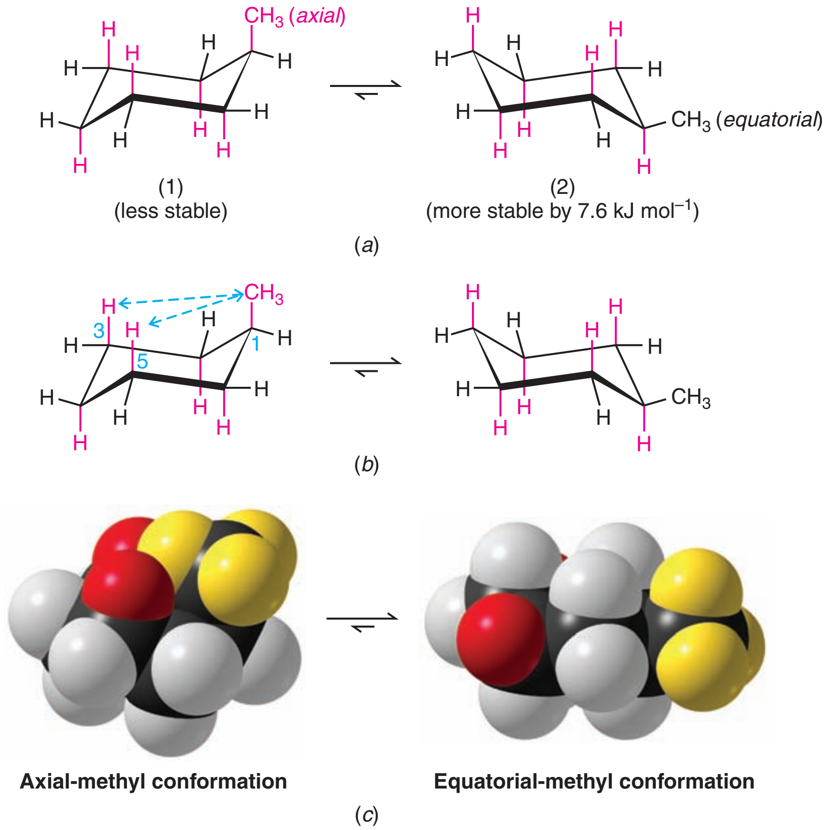


Figure 4.19 *(a)* The conformations of methylcyclohexane with the methyl group axial (1) and equatorial (2). *(b)* **1,3-Diaxial interactions** between the two axial hydrogen atoms and the axial methyl group in the axial conformation of methylcyclohexane are shown with dashed arrows. Less crowding occurs in the equatorial conformation. *(c)* Space-filling molecular models for the axial–methyl and equatorial–methyl conformers of methylcyclohexane. In the axial–methyl conformer the methyl group (shown with yellow hydrogen atoms) is crowded by the 1,3-diaxial hydrogen atoms (red), as compared to the equatorial–methyl conformer, which has no 1,3-diaxial interactions with the methyl group.

hydrogens) occupies an *axial* position, and in the other the methyl group occupies an *equatorial* position.

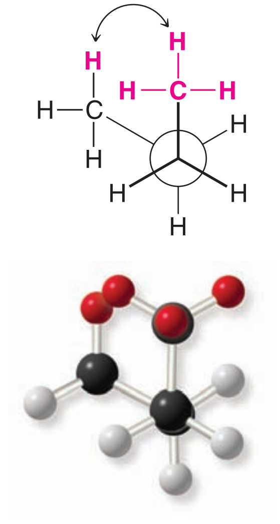
- The most stable conformation for a monosubstituted cyclohexane ring (a cyclohexane ring where one carbon atom bears a group other than hydrogen) is the conformation where the substituent is equatorial.

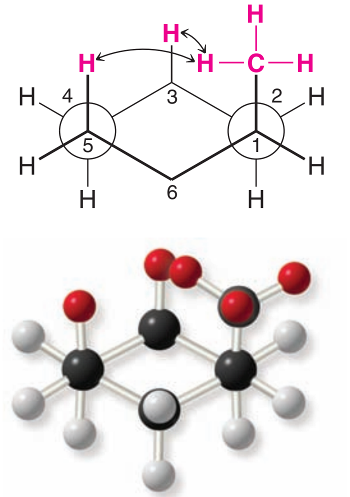
Studies indicate that the conformation with the equatorial methyl group is more stable than the conformation with the axial methyl group by about 7.6 kJ mol^{-1}. Thus, in the equilibrium mixture, the conformation with the methyl group in the equatorial position is the predominant one, constituting about 95% of the equilibrium mixture.

The greater stability of methylcyclohexane with an equatorial methyl group can be understood through an inspection of the two forms as they are shown in Figs. 4.19*a–c*.

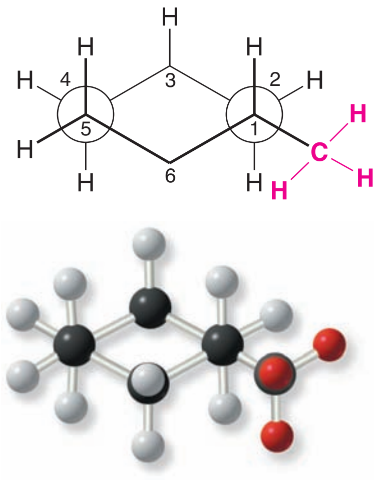
- Studies done with models of the two conformations show that when the methyl group is axial, it is so close to the two axial hydrogens on the same side of the ring (attached to the C3 and C5 atoms) that **the dispersion forces between them are repulsive**.
- This type of steric strain, because it arises from an interaction between an axial group on carbon atom 1 and an axial hydrogen on carbon atom 3 (or 5) is called a **1,3-diaxial interaction**.
- Studies with other substituents show that **there is generally less repulsion when any group larger than hydrogen is equatorial rather than axial**.

The strain caused by a 1,3-diaxial interaction in methylcyclohexane is the same as the strain caused by the close proximity of the hydrogen atoms of methyl groups in the gauche form of butane (Section 4.9). Recall that the interaction in *gauche*-butane (called, for convenience, a *gauche interaction*) causes *gauche*-butane to be less stable than *anti*-butane by 3.8 kJ mol^{-1}. The following Newman projections will help you to see that the two steric interactions are the same. In the second projection we view axial methylcyclohexane along the C1—C2 bond and see that what we call a 1,3-diaxial interaction is simply a gauche interaction between the hydrogen atoms of the methyl group and the hydrogen atom at C3:


gauche-Butane
(3.8 kJ mol^{-1} steric strain)



Axial methylcyclohexane
(two gauche interactions =
7.6 kJ mol^{-1} steric strain)



Equatorial methylcyclohexane
(more stable by 7.6 kJ mol^{-1})

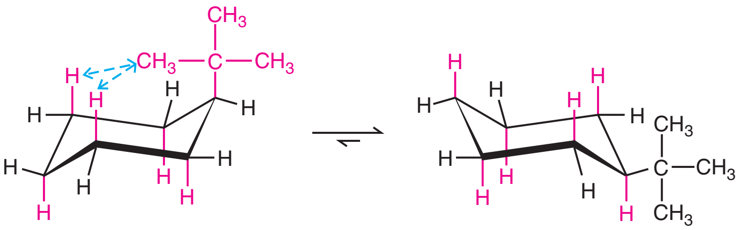
Viewing methylcyclohexane along the C1—C6 bond (do this with a model) shows that it has a second identical gauche interaction between the hydrogen atoms of the methyl group and the hydrogen atom at C5. The methyl group of axial methylcyclohexane, therefore, has two gauche interactions and, consequently, it has 7.6 kJ mol^{-1} of strain. The methyl group of equatorial methylcyclohexane does not have a gauche interaction because it is anti to C3 and C5.

Review Problem 4.14

Show by a calculation (using the formula $\Delta G^\circ = -RT \ln K_{eq}$) that a free-energy difference of 7.6 kJ mol^{-1} between the axial and equatorial forms of methylcyclohexane at 25°C (with the equatorial form being more stable) does correlate with an equilibrium mixture in which the concentration of the equatorial form is approximately 95%.

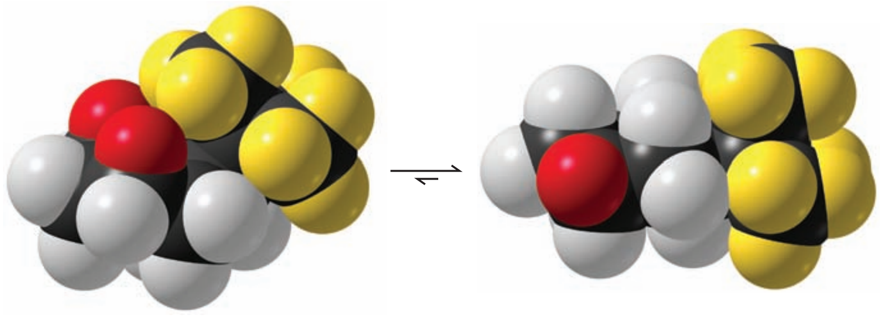
4.12C 1,3-Diaxial Interactions of a *tert*-Butyl Group

In cyclohexane derivatives with larger alkyl substituents, the strain caused by 1,3-diaxial interactions is even more pronounced. The conformation of *tert*-butylcyclohexane with the *tert*-butyl group equatorial is estimated to be approximately 21 kJ mol^{-1} more stable than the axial form (Fig. 4.20). This large energy difference between the two conformations means that, at room temperature, 99.99% of the molecules of *tert*-butylcyclohexane have the *tert*-butyl group in the equatorial position. (The molecule is not conformationally "locked," however; it still flips from one chair conformation to the other.)


Axial *tert*-butylcyclohexane **Equatorial *tert*-butylcyclohexane**

(*a*)

Figure 4.20 *(a)* Diaxial interactions with the large *tert*-butyl group axial cause the conformation with the *tert*-butyl group equatorial to be the predominant one to the extent of 99.99%. *(b)* Space-filling molecular models of *tert*-butylcyclohexane in the axial (ax) and equatorial (eq) conformations, highlighting the position of the 1,3-hydrogens (red) and the *tert*-butyl group (shown with yellow hydrogen atoms).

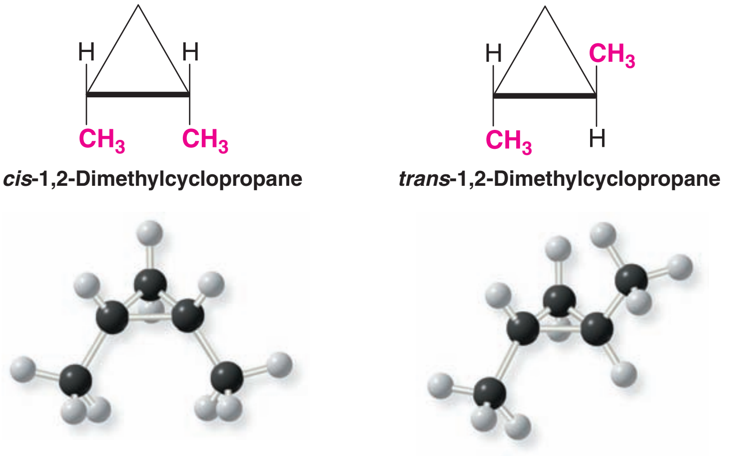

(ax) **(eq)**

(*b*)

Figure 4.20 (*continued*)

4.13 Disubstituted Cycloalkanes: Cis–Trans Isomerism

The presence of two substituents on different carbons of a cycloalkane allows for the possibility of **cis–trans isomerism** similar to the kind we saw for alkenes in Section 1.13B. These cis–trans isomers are also **stereoisomers** because they differ from each other only in the arrangement of their atoms in space. Consider 1,2-dimethylcyclopropane (Fig. 4.21) as an example.



***cis*-1,2-Dimethylcyclopropane** ***trans*-1,2-Dimethylcyclopropane**

Figure 4.21 The *cis-* and *trans-*1,2-dimethylcyclopropane isomers.

The planarity of the cyclopropane ring makes the cis–trans isomerism obvious. In the first structure the methyl groups are on the same side of the ring; therefore, they are cis. In the second structure, they are on opposite sides of the ring; they are trans.

Cis and trans isomers such as these cannot be interconverted without breaking carbon–carbon bonds. They will have different physical properties (boiling points, melting points, and so on). As a result, they can be separated, placed in separate bottles, and kept indefinitely.

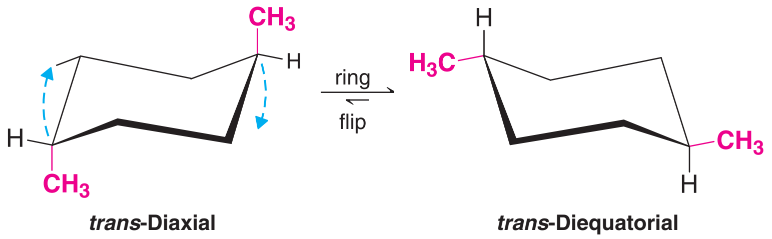
Review Problem 4.15

Write structures for the cis and trans isomers of **(a)** 1,2-dichlorocyclopentane and **(b)** 1,3-dibromocyclobutane. **(c)** Are cis–trans isomers possible for 1,1-dibromocyclobutane?

4.13A Cis–Trans Isomerism and Conformational Structures of Cyclohexanes

Trans 1,4-Disubstituted Cyclohexanes If we consider dimethylcyclohexanes, the structures are somewhat more complex because the cyclohexane ring is not planar. Beginning with *trans*-1,4-dimethylcyclohexane, because it is easiest to visualize, we find there are two possible chair conformations (Fig. 4.22). In one conformation both methyl groups are axial; in the other both are equatorial. The diequatorial conformation is, as we would expect it to be, the more stable conformation, and it represents the structure of at least 99% of the molecules at equilibrium.

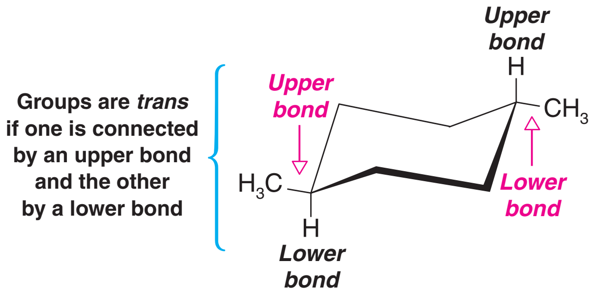
That the diaxial form of *trans*-1,4-dimethylcyclohexane is a trans isomer is easy to see; the two methyl groups are clearly on opposite sides of the ring. The trans relationship of the methyl groups in the diequatorial form is not as obvious, however.



***trans*-Diaxial** ***trans*-Diequatorial**

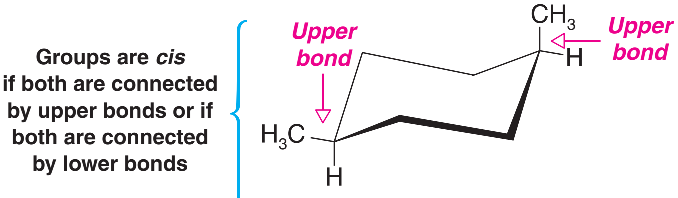
Figure 4.22 The two chair conformations of *trans*-1,4-dimethylcyclohexane: trans–diequatorial and trans–diaxial. The trans–diequatorial form is more stable by 15.2 kJ mol^{-1}.

How do we know two groups are cis or trans? A general way to recognize a trans-disubstituted cyclohexane is to notice that one group is attached by the *upper* bond (of the two to its carbon) and one by the *lower* bond:


Groups are *trans* if one is connected by an upper bond and the other by a lower bond

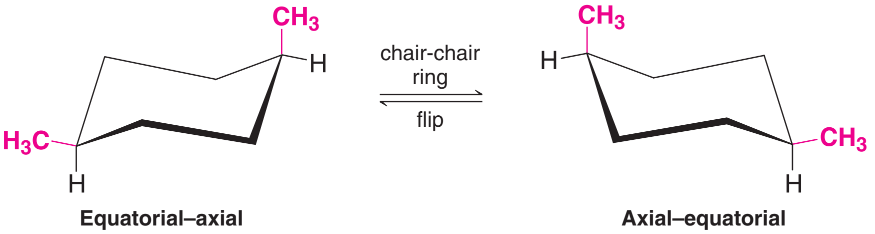
***trans*-1,4-Dimethylcyclohexane**

In a cis 1,4-disubstituted cyclohexane both groups are attached by an upper bond or both by a lower bond. For example,


Groups are *cis* if both are connected by upper bonds or if both are connected by lower bonds

***cis*-1,4-Dimethylcyclohexane**

Cis 1,4-Disubstituted Cyclohexanes *cis*-1,4-Dimethylcyclohexane exists in two *equivalent* chair conformations (Fig. 4.23). In a cis 1,4-disubstituted cyclohexane, one group is axial and the other is equatorial in both of the possible chair conformations.

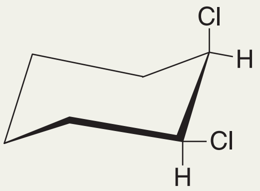


Equatorial–axial **Axial–equatorial**

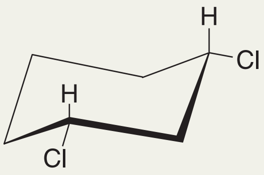
Figure 4.23 Equivalent conformations of *cis*-1,4-dimethylcyclohexane.

Solved Problem 4.5

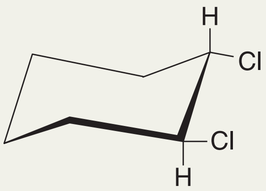
Consider each of the following conformational structures and tell whether each is cis or trans:



(a)



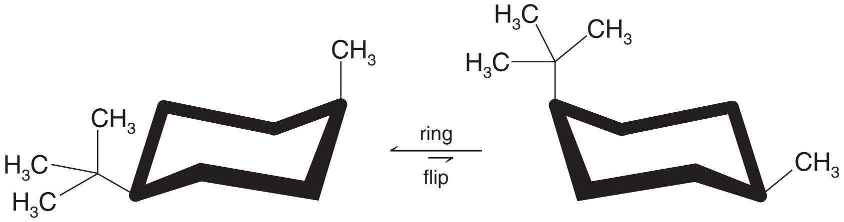
(b)



(c)

ANSWER **(a)** Each chlorine is attached by the upper bond at its carbon; therefore, both chlorine atoms are on the same side of the molecule and this is a cis isomer. This is a *cis*-1,2-dichlorocyclohexane. **(b)** Here both chlorine atoms are attached by a lower bond; therefore, in this example, too, both chlorine atoms are on the same side of the molecule and this, too, is a cis isomer. It is *cis*-1,3-dichlorocyclohexane. **(c)** Here one chlorine atom is attached by a lower bond and one by an upper bond. The two chlorine atoms, therefore, are on opposite sides of the molecule, and this is a trans isomer. It is *trans*-1,2-dichlorocyclohexane. Verify these facts by building models.

The two conformations of cis 1,4-disubstituted cyclohexanes *are not equivalent* if one group is larger than the other. Consider *cis*-1-*tert*-butyl-4-methylcyclohexane:



(More stable because large group is equatorial)

(Less stable because large group is axial)

***cis*-1-*tert*-Butyl-4-methylcyclohexane**

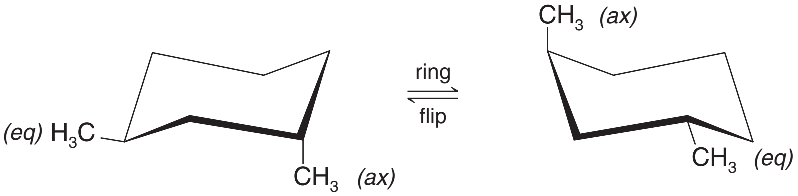
Here the more stable conformation is the one with the larger group equatorial. This is a general principle:

- When one ring substituent group is larger than the other and they cannot both be equatorial, the conformation with the larger group equatorial will be more stable.

Review Problem 4.16

(a) Write structural formulas for the two chair conformations of *cis*-1-isopropyl-4-methylcyclohexane. **(b)** Are these two conformations equivalent? **(c)** If not, which would be more stable? **(d)** Which would be the preferred conformation at equilibrium?

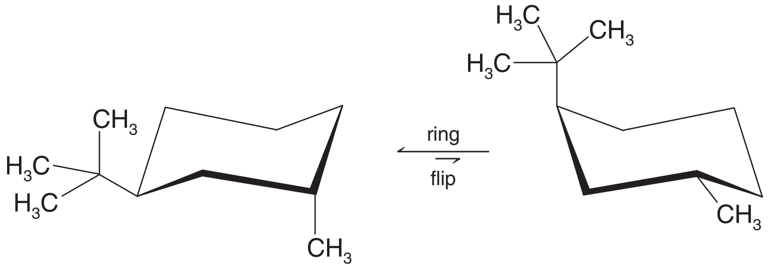
Trans 1,3-Disubstituted Cyclohexanes *trans*-1,3-Dimethylcyclohexane is like the cis 1,4 compound in that each conformation has one methyl group in an axial position and one methyl group in an equatorial position. The following two conformations are of equal energy and are equally populated at equilibrium:



***trans*-1,3-Dimethylcyclohexane**

Equal energy and equally populated conformations

The situation is different for *trans*-1-*tert*-butyl-3-methylcyclohexane (shown below) because the two ring substituents are not the same. Again, we find that the lower energy conformation is that with the largest group equatorial.

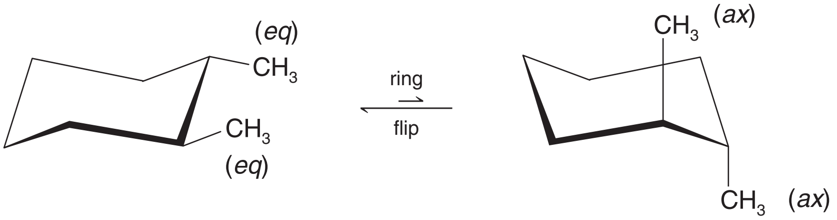


(More stable because large group is equatorial) **(Less stable because large group is axial)**

***trans*-1-*tert*-Butyl-3-methylcyclohexane**

Cis 1,3-Disubstituted Cyclohexanes *cis*-1,3-Dimethylcyclohexane has a conformation in which both methyl groups are equatorial and one in which both methyl groups are axial. **As we would expect, the conformation with both methyl groups equatorial is the more stable one**.

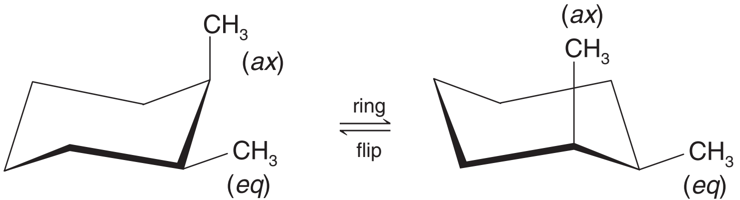
Trans 1,2-Disubstituted Cyclohexanes *trans*-1,2-Dimethylcyclohexane has a conformation in which both methyl groups are equatorial and one in which both methyl groups are axial. **As we would expect, the conformation with both methyl groups equatorial is the more stable one**.



(Diequatorial is much more stable) **(Diaxial is much less stable)**

***trans*-1,2-Dimethylcyclohexane**

Cis 1,2-Disubstituted Cyclohexanes *cis*-1,2-Dimethylcyclohexane has one methyl group that is axial and one methyl group that is equatorial in each of its chair conformations, thus its two conformations are of equal stability.



(Equatorial-axial) **(Axial-equatorial)**

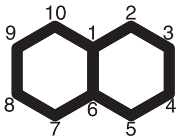
***cis*-1,2-Dimethylcyclohexane**
Equal energy and equally populated conformations

Review Problem 4.17 Write a conformational structure for 1-bromo-3-chloro-5-fluorocyclohexane in which all the substituents are equatorial. Then write its structure after a ring flip.

Review Problem 4.18 **(a)** Write the two conformations of *cis*-1-*tert*-butyl-2-methylcyclohexane. **(b)** Which conformer has the lowest potential energy?

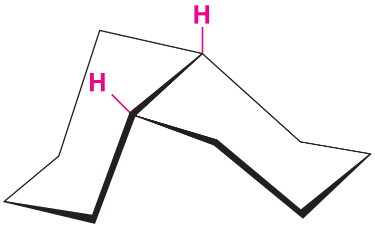
4.14 *Bicyclic and Polycyclic Alkanes*

Many of the molecules that we encounter in our study of organic chemistry contain more than one ring (Section 4.4B). One of the most important bicyclic systems is bicyclo[4.4.0]decane, a compound that is usually called by its common name, *decalin*:

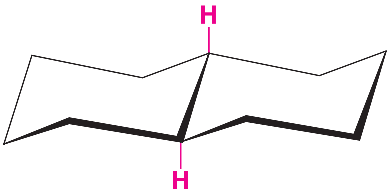


Decalin (bicyclo[4.4.0]decane)
(carbon atoms 1 and 6 are bridgehead carbon atoms)

Decalin shows cis–trans isomerism:



***cis*-Decalin**

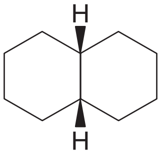


***trans*-Decalin**

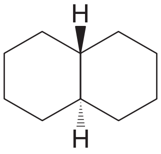
Helpful Hint

Chemical Abstracts Service (CAS) determines the number of rings by the formula $S - A + 1 = N$, where S is the number of single bonds in the ring system, A is the number of atoms in the ring system, and N is the calculated number of rings (see Problem 4.30).

In *cis*-decalin the two hydrogen atoms attached to the bridgehead atoms lie on the same side of the ring; in *trans*-decalin they are on opposite sides. We often indicate this by writing their structures in the following way:



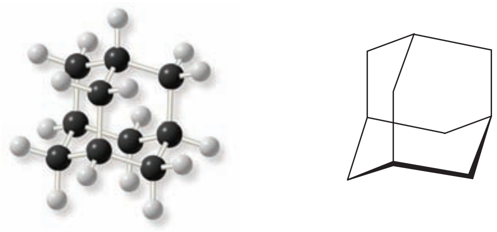
***cis*-Decalin**



***trans*-Decalin**

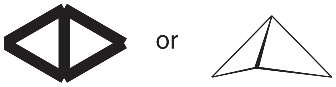
Simple rotations of groups about carbon–carbon bonds do not interconvert *cis*- and *trans*-decalins. They are stereoisomers and they have different physical properties.

Adamantane is a tricyclic system that contains a three-dimensional array of cyclohexane rings, all of which are in the chair form.

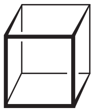


Adamantane

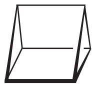
One goal of research in recent years has been the synthesis of unusual, and sometimes highly strained, cyclic hydrocarbons. Among those that have been prepared are the compounds that follow:



Bicyclo[1.1.0]butane



Cubane



Prismane

THE CHEMISTRY OF . . .

Elemental Carbon

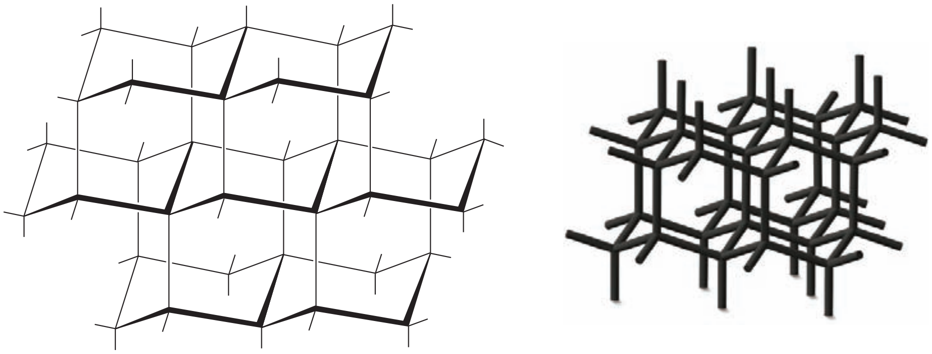
Carbon as the pure element exists in several forms that are as different from one another as it is possible to imagine. Different forms of a pure element are called allotropes. One allotrope of pure carbon is the very soft and totally black substance called **graphite**, the main substance at the center of pencils and the main component of charcoal and chimney soot. Another allotropic form of carbon is **diamond**, the colorless brilliant gem that is the hardest of all substances found in nature. Still another allotrope, perhaps the most exotic, is called **buckminsterfullerene** after the inventor of the geodesic dome, Buckminster Fuller.

The different properties of these allotropic forms arise from different structural arrangements of the carbon atoms in each form, and these arrangements result, in part, from different hybridization states of their carbon atoms. The carbon atoms of diamond are all sp^3 hybridized with tetrahedrally oriented bonds. The structure of diamond is what you would get if you extended the structure of adamantane in three dimensions. The great hardness of diamond results from the fact that the entire diamond crystal is one large molecule—a network of interconnecting rings that is held together by millions of strong covalent bonds.

In graphite the carbon atoms are sp^2 hybridized. Because of the trigonal planar orientation of their covalent bonds, the carbon atoms of graphite are in sheets. The sheets are actually huge molecules consisting of fused benzene rings (see below). While all of the covalent bonds of each sheet lie in the same plane, the sheets are piled one on another and the *p* orbitals of their benzene rings keep them apart. Although these *p* orbitals interact, their interactions are very weak, much weaker than those of covalent bonds, allowing the individual sheets to slide past one another and accounting for graphite's usefulness as a lubricant.

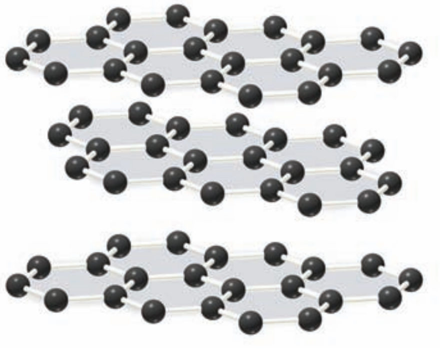
Buckminsterfullerene (shown on the next page) is a representative of a new class of carbon compounds discovered in 1985 consisting of carbon clusters called fullerenes (see Section 14.8C for the story of their discovery and synthesis). Buckminsterfullerene (also called a "buckyball") is a hollow cluster of 60 carbon atoms, all of which are sp^2 hybridized, and which are joined together in a pattern like the seams of a soccer ball. The center of the buckyball is large enough to hold an atom of helium or argon, and such compounds are known. In the buckyball there are 32 interlocking rings: 20 are hexagons and 12 are pentagons, producing a highly symmetrical molecule. A smaller symmetrical molecule, synthesized in 1982 by Leo A. Paquette and co-workers at Ohio State University, is dodecahedrane.

One final point: We began this book telling of how all of the carbon atoms of the universe are thought to have been formed in the interiors of stars and to have been dispersed



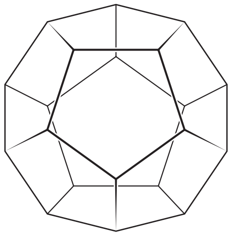
A portion of the diamond structure

Carbon is shown here in its diamond and graphite forms

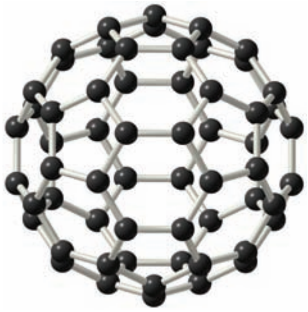


A portion of the structure of graphite

throughout the universe when some of those stars exploded as supernovae. Consider this evidence. Sediments on our planet, known to be 251 million years old and which were formed at the time of a great extinction caused by the collision of a comet with Earth, have been found to contain buckyballs with helium atoms in their centers. The isotopic ratio of $^3He/^4He$ in them is much larger than the ratio in ordinary helium found on Earth now, indicating that the helium was of extraterrestrial origin. So in these discoveries we have fascinating evidence for the origin of elemental carbon and how some of it got here. Most carbon atoms were produced when Earth was formed billions of years ago. But the carbon atoms of the buckyballs found in this sediment, formed originally in the interior of a star somewhere in the universe, probably made their way here 251 million years ago in a comet or meteorite.



Dodecahedrane



Buckminsterfullerene

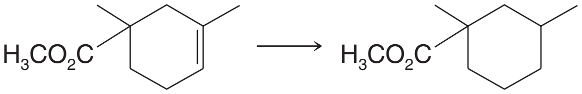
4.15 *Chemical Reactions of Alkanes*

Alkanes, as a class, are characterized by a general inertness to many chemical reagents. Carbon–carbon and carbon–hydrogen bonds are quite strong; they do not break unless alkanes are heated to very high temperatures. Because carbon and hydrogen atoms have nearly the same electronegativity, the carbon–hydrogen bonds of alkanes are only slightly polarized. As a consequence, they are generally unaffected by most bases. Molecules of alkanes have no unshared electrons to offer as sites for attack by acids. This low reactivity of alkanes toward many reagents accounts for the fact that alkanes were originally called **paraffins** (*parum affinis*, Latin: little affinity).

The term paraffin, however, was probably not an appropriate one. We all know that alkanes react vigorously with oxygen when an appropriate mixture is ignited. This combustion occurs, for example, in the cylinders of automobiles, in furnaces, and, more gently, with paraffin candles. When heated, alkanes also react with chlorine and bromine, and they react explosively with fluorine. We shall study these reactions in Chapter 10.

4.16 *Synthesis of Alkanes and Cycloalkanes*

A chemical synthesis may require, at some point, the conversion of a carbon–carbon double or triple bond to a single bond. Synthesis of the following compound, used as an ingredient in some perfumes, is an example.



(used in some perfumes)

This conversion is easily accomplished by a reaction called **hydrogenation**. There are several reaction conditions that can be used to carry out hydrogenation, but among the common ways is use of hydrogen gas and a solid metal catalyst such as platinum, palladium, or nickel. Equations in the following section represent general examples for the hydrogenation of alkenes and alkynes.

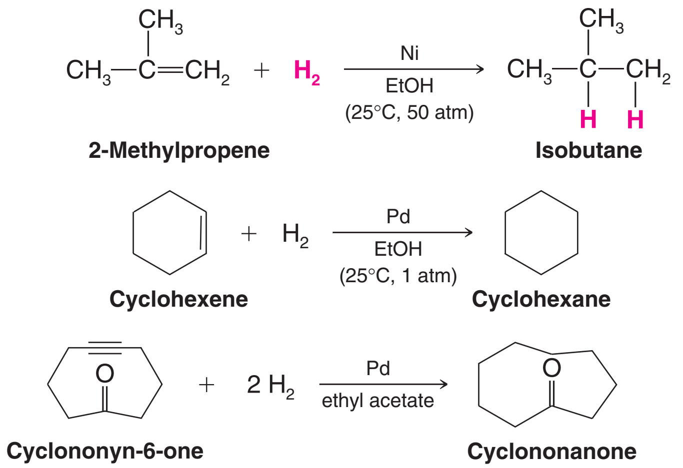
4.16A Hydrogenation of Alkenes and Alkynes

Alkenes and alkynes react with hydrogen in the presence of metal catalysts such as nickel, palladium, and platinum to produce alkanes. The general reaction is one in which the atoms of the hydrogen molecule add to each atom of the carbon–carbon double or triple bond of the alkene or alkyne. This converts the alkene or alkyne to an alkane:

General Reaction

The reaction is usually carried out by dissolving the alkene or alkyne in a solvent such as ethyl alcohol (C_2H_5OH), adding the metal catalyst, and then exposing the mixture to hydrogen gas under pressure in a special apparatus. One molar equivalent of hydrogen is required to reduce an alkene to an alkane. Two molar equivalents are required to reduce an alkyne. (We shall discuss the mechanism of this reaction in Chapter 7.)

Specific Examples



Review Problem 4.19 Show the reactions involved for hydrogenation of all the alkenes and alkynes that would yield 2-methylbutane.

4.17 How to Gain Structural Information from Molecular Formulas and the Index of Hydrogen Deficiency

A chemist working with an unknown compound can obtain considerable information about its structure from the compound's molecular formula and its **index of hydrogen deficiency (IHD)**.

- The **index of hydrogen deficiency (IHD)*** is defined as the difference in the *number of pairs* of hydrogen atoms between the compound under study and an acyclic alkane having the same number of carbons.

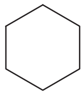
Saturated acyclic hydrocarbons have the general molecular formula C_nH_{2n+2}. Each double bond or ring reduces the number of hydrogen atoms by two as compared with the formula for a saturated compound. Thus each ring or double bond provides one unit of

---

*Some organic chemists refer to the index of hydrogen deficiency as the "degree of unsaturation" or "the number of double-bond equivalencies."

hydrogen deficiency. For example, 1-hexene and cyclohexane have the same molecular formula (C_6H_{12}) and they are constitutional isomers.

1-Hexene
(C_6H_{12})

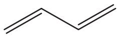


Cyclohexane
(C_6H_{12})

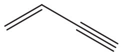
Both 1-hexene and cyclohexane (C_6H_{12}) have an index of hydrogen deficiency equal to 1 (meaning one pair of hydrogen atoms), because the corresponding acyclic alkane is hexane (C_6H_{14}).

C_6H_{14} = formula of corresponding alkane (hexane)
$\underline{C_6H_{12}}$ = formula of compound (1-hexene or cyclohexane)
H_2 = difference = 1 pair of hydrogen atoms
Index of hydrogen deficiency = 1

Alkynes and alkadienes (alkenes with two double bonds) have the general formula C_nH_{2n-2}. Alkenynes (hydrocarbons with one double bond and one triple bond) and alkatrienes (alkenes with three double bonds) have the general formula C_nH_{2n-4}, and so forth.



1,3-Butadiene
IHD = 2



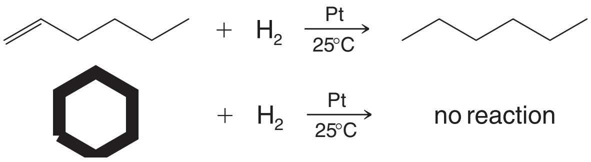
But-1-en-3-yne
IHD = 3

1,3,5-Hexatriene
IHD = 3

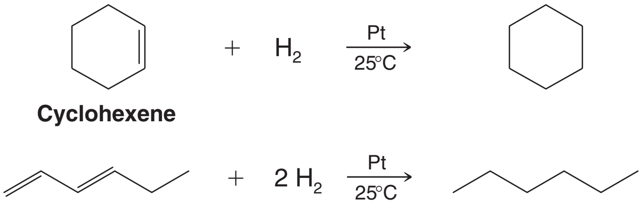
The index of hydrogen deficiency is easily determined by comparing the molecular formula of a given compound with the formula for its hydrogenation product.

- Each double bond consumes one molar equivalent of hydrogen and counts for one unit of hydrogen deficiency.
- Each triple bond consumes two molar equivalents of hydrogen and counts for two units of hydrogen deficiency.
- Rings are not affected by hydrogenation, but each ring still counts for one unit of hydrogen deficiency.

Hydrogenation, therefore, allows us to distinguish between rings and double or triple bonds. Consider again two compounds with the molecular formula C_6H_{12}: 1-hexene and cyclohexane. 1-Hexene reacts with one molar equivalent of hydrogen to yield hexane; under the same conditions cyclohexane does not react:



Or consider another example. Cyclohexene and 1,3-hexadiene have the same molecular formula (C_6H_{10}). Both compounds react with hydrogen in the presence of a catalyst, but cyclohexene, because it has a ring and only one double bond, reacts with only one molar equivalent. 1,3-Hexadiene adds two molar equivalents:



Cyclohexene

1,3-Hexadiene

Review Problem 4.20

(a) What is the index of hydrogen deficiency of 2-hexene? **(b)** Of methylcyclopentane? **(c)** Does the index of hydrogen deficiency reveal anything about the location of the double bond in the chain? **(d)** About the size of the ring? **(e)** What is the index of hydrogen deficiency of 2-hexyne? **(f)** In general terms, what structural possibilities exist for a compound with the molecular formula $C_{10}H_{16}$?

Review Problem 4.21

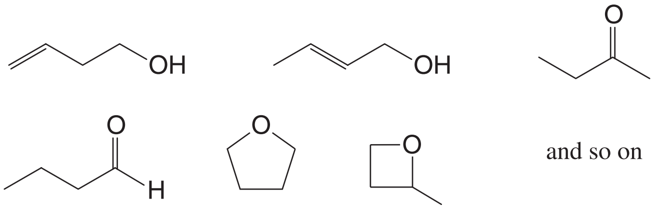
Zingiberene, a fragrant compound isolated from ginger, has the molecular formula $C_{15}H_{24}$ and is known not to contain any triple bonds. **(a)** What is the index of hydrogen deficiency of zingiberene? **(b)** When zingiberene is subjected to catalytic hydrogenation using an excess of hydrogen, 1 mol of zingiberene absorbs 3 mol of hydrogen and produces a compound with the formula $C_{15}H_{30}$. How many double bonds does a molecule of zingiberene have? **(c)** How many rings?

4.17A Compounds Containing Halogens, Oxygen, or Nitrogen

Calculating the index of hydrogen deficiency (IHD) for compounds other than hydrocarbons is relatively easy.

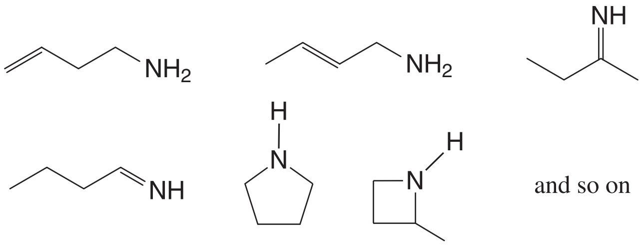
For compounds containing halogen atoms, we simply count the halogen atoms as though they were hydrogen atoms. Consider a compound with the formula $C_4H_6Cl_2$. To calculate the IHD, we change the two chlorine atoms to hydrogen atoms, considering the formula as though it were C_4H_8. This formula has two hydrogen atoms fewer than the formula for a saturated alkane (C_4H_{10}), and this tells us that the compound has IHD = 1. It could, therefore, have either one ring or one double bond. [We can tell which it has from a hydrogenation experiment: If the compound adds one molar equivalent of hydrogen (H_2) on catalytic hydrogenation at room temperature, then it must have a double bond; if it does not add hydrogen, then it must have a ring.]

For compounds containing oxygen, we simply ignore the oxygen atoms and calculate the IHD from the remainder of the formula. Consider as an example a compound with the formula C_4H_8O. For the purposes of our calculation we consider the compound to be simply C_4H_8 and we calculate IHD = 1. Again, this means that the compound contains either a ring or a double bond. Some structural possibilities for this compound are shown next. Notice that the double bond may be present as a carbon–oxygen double bond:



and so on

For compounds containing nitrogen atoms we subtract one hydrogen for each nitrogen atom, and then we ignore the nitrogen atoms. For example, we treat a compound with the formula C_4H_9N as though it were C_4H_8, and again we get IHD = 1. Some structural possibilities are the following:



and so on

Review Problem 4.22

Carbonyl groups also count for a unit of hydrogen deficiency. What are the indices of hydrogen deficiency for the reactant and for the product in the equation shown at the beginning of Section 4.16 for synthesis of a perfume ingredient?

4.18 Applications of Basic Principles

In this chapter we have seen repeated applications of one basic principle in particular:

Nature Prefers States of Lower Potential Energy This principle underlies our explanations of conformational analysis in Sections 4.8–4.13. The staggered conformation of ethane (Section 4.8) is preferred (more populated) in a sample of ethane because its potential energy is lowest. In the same way, the anti conformation of butane (Section 4.9) and the chair conformation of cyclohexane (Section 4.11) are the preferred conformations of these molecules because these conformations are of lowest potential energy. Methylcyclohexane (Section 4.12) exists mainly in the chair conformation with its methyl group equatorial for the same reason. Disubstituted cycloalkanes (Section 4.13) prefer a conformation with both substituents equatorial if this is possible, and, if not, they prefer a conformation with the larger group equatorial. The preferred conformation in each instance is the one of lowest potential energy.

Another effect that we encounter in this chapter, and one we shall see again and again, is how **steric factors** (spatial factors) can affect the stability and reactivity of molecules. Unfavorable spatial interactions between groups are central to explaining why certain conformations are higher in energy than others. But fundamentally this effect is derived itself from another familiar principle: **like charges repel.** Repulsive interactions between the electrons of groups that are in close proximity cause certain conformations to have higher potential energy than others. We call this kind of effect *steric hindrance*.

In This Chapter

One of the reasons we organic chemists love our discipline is that, besides knowing each molecule has a family, we also know that each one has its own architecture, "personality," and unique name. You have already learned in Chapters 1–3 about molecular personalities with regard to charge distribution, polarity, and relative acidity or basicity. In this chapter you have now learned how to give unique names to simple molecules using the IUPAC system. You also learned more about the overall shapes of organic molecules, how their shapes can change through bond rotations, and how we can compare the relative energies of those changes using conformational analysis. You now know that the extent of flexibility or rigidity in a molecule has to do with the types of bonds present (single, double, triple), and whether there are rings or bulky groups that inhibit bond rotation. Some organic molecules are very flexible members of the family, such as the molecules in our muscle fibers, while others are very rigid, like the carbon lattice of diamond. Most molecules, however, have both flexible and rigid aspects to their structures. With the knowledge from this chapter, added to other fundamentals you have already learned, you are on your way to developing an understanding of organic chemistry that we hope will be as strong as diamonds, and that you can flex like a muscle when you approach a problem. When you are finished with this chapter's homework, maybe you can even take a break by resting your mind on the chair conformation of cyclohexane.

Key Terms and Concepts

The key terms and concepts that are highlighted in **bold, blue text** within the chapter are defined in the glossary (at the back of the book) and have hyperlinked definitions in the accompanying ***WileyPLUS*** course (www.wileyplus.com).

Problems

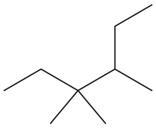
Note to Instructors: Many of the homework problems are available for assignment via *WileyPLUS*, an online teaching and learning solution program.

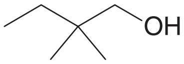
NOMENCLATURE AND ISOMERISM

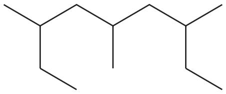
4.23 Write a bond-line formula for each of the following compounds:

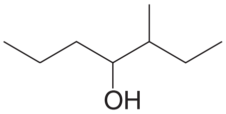
(a) 1,4-Dichloropentane	**(f)** 1,1-Dichlorocyclopentane	**(k)** 1,4-Dicyclopropylhexane
(b) *sec*-Butyl bromide	**(g)** *cis*-1,2-Dimethylcyclopropane	**(l)** Neopentyl alcohol
(c) 4-Isopropylheptane	**(h)** *trans*-1,2-Dimethylcyclopropane	**(m)** Bicyclo[2.2.2]octane
(d) 2,2,3-Trimethylpentane	**(i)** 4-Methyl-2-pentanol	**(n)** Bicyclo[3.1.1]heptane
(e) 3-Ethyl-2-methylhexane	**(j)** *trans*-4-Isobutylcyclohexanol	**(o)** Cyclopentylcyclopentane

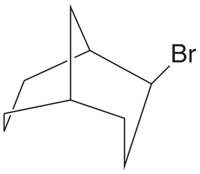
4.24 Give systematic IUPAC names for each of the following:

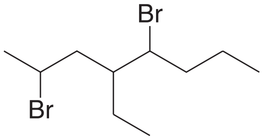
(a)


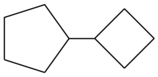
(b)


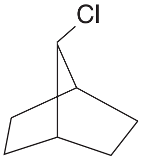
(c)


(d)


(e)


(f)


(g)


(h)


4.25 The name *sec*-butyl alcohol defines a specific structure but the name *sec*-pentyl alcohol is ambiguous. Explain.

4.26 Write the structure and give the IUPAC systematic name of an alkane or cycloalkane with the formulas **(a)** C_8H_{18} that has only primary hydrogen atoms, **(b)** C_6H_{12} that has only secondary hydrogen atoms, **(c)** C_6H_{12} that has only primary and secondary hydrogen atoms, and **(d)** C_8H_{14} that has 12 secondary and 2 tertiary hydrogen atoms.

4.27 Write the structure(s) of the simplest alkane(s), i.e., one(s) with the fewest number of carbon atoms, wherein each possesses primary, secondary, tertiary, and quaternary carbon atoms. (A quaternary carbon is one that is bonded to four other carbon atoms.) Assign an IUPAC name to each structure.

4.28 Ignoring compounds with double bonds, write structural formulas and give names for all of the isomers with the formula C_5H_{10}.

4.29 Write structures for the following bicyclic alkanes:

(a) Bicyclo[1.1.0]butane	**(c)** 2-Chlorobicyclo[3.2.0]heptane
(b) Bicyclo[2.1.0]pentane	**(d)** 7-Methylbicyclo[2.2.1]heptane

4.30 Use the $S - A + 1 = N$ method (Helpful Hint, Section 4.14) to determine the number of rings in cubane (Section 4.14).

4.31 A spiro ring junction is one where two rings that share no bonds originate from a single carbon atom. Alkanes containing such a ring junction are called spiranes.

(a) For the case of bicyclic spiranes of formula C_7H_{12}, write structures for all possibilities where all carbons are incorporated into rings.

(b) Write structures for other bicyclic molecules that fit this formula.

4.32 Tell what is meant by an homologous series and illustrate your answer by writing structures for an homologous series of alkyl halides.

HYDROGENATION

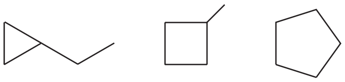
4.33 Four different cycloalkenes will all yield methylcyclopentane when subjected to catalytic hydrogenation. What are their structures? Show the reactions.

4.34 **(a)** Three different alkenes yield 2-methylbutane when they are hydrogenated in the presence of a metal catalyst. Give their structural formulas and write equations for the reactions involved. **(b)** One of these alkene isomers has characteristic absorptions at approximately 998 and 914 cm^{-1} in its IR spectrum. Which one is it?

4.35 An alkane with the formula C_6H_{14} can be prepared by hydrogenation of either of two precursor alkenes having the formula C_6H_{12}. Write the structure of this alkane, give its IUPAC name, and show the reactions.

CONFORMATIONS AND STABILITY

4.36 Rank the following compounds in order of increasing stability based on relative ring strain.



4.37 Write the structures of two chair conformations of 1-*tert*-butyl-1-methylcyclohexane. Which conformation is more stable? Explain your answer.

4.38 Sketch curves similar to the one given in Fig. 4.8 showing the energy changes that arise from rotation about the C2—C3 bond of **(a)** 2,3-dimethylbutane and **(b)** 2,2,3,3-tetramethylbutane. You need not concern yourself with actual numerical values of the energy changes, but you should label all maxima and minima with the appropriate conformations.

4.39 Without referring to tables, decide which member of each of the following pairs would have the higher boiling point. Explain your answers.

(**a**) Pentane or 2-methylbutane
(**b**) Heptane or pentane
(**c**) Propane or 2-chloropropane
(**d**) Butane or 1-propanol
(**e**) Butane or CH_3COCH_3

4.40 One compound whose molecular formula is C_4H_6 is a bicyclic compound. Another compound with the same formula has an infrared absorption at roughly 2250 cm^{-1} (the bicyclic compound does not). Draw structures for each of these two compounds and explain how the IR absorption allows them to be differentiated.

4.41 Which compound would you expect to be the more stable: *cis*-1,2-dimethylcyclopropane or *trans*-1,2-dimethylcyclopropane? Explain your answer.

4.42 Consider that cyclobutane exhibits a puckered geometry. Judge the relative stabilities of the 1,2-disubstituted cyclobutanes and of the 1,3-disubstituted cyclobutanes. (You may find it helpful to build handheld molecular models of representative compounds.)

4.43 Write the two chair conformations of each of the following and in each part designate which conformation would be the more stable: **(a)** *cis*-1-*tert*-butyl-3-methylcyclohexane, **(b)** *trans*-1-*tert*-butyl-3-methylcyclohexane, **(c)** *trans*-1-*tert*-butyl-4-methylcyclohexane, **(d)** *cis*-1-*tert*-butyl-4-methylcyclohexane.

4.44 Provide an explanation for the surprising fact that all-*trans*-1,2,3,4,5,6-hexaisopropylcyclohexane is a stable molecule in which all isopropyl groups are axial. (You may find it helpful to build a handheld molecular model.)

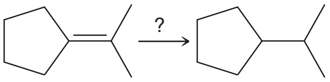
4.45 *trans*-1,3-Dibromocyclobutane has a measurable dipole moment. Explain how this proves that the cyclobutane ring is not planar.

SYNTHESIS

4.46 Specify the missing compounds and/or reagents in each of the following syntheses:

(**a**) *trans*-5-Methyl-2-hexene $\xrightarrow{?}$ 2-methylhexane

(**b**)



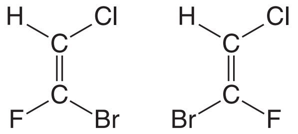
(**c**) Chemical reactions rarely yield products in such initially pure form that no trace can be found of the starting materials used to make them. What evidence in an IR spectrum of each of the crude (unpurified) products from the above reactions would indicate the presence of one of the organic reactants used to synthesize each target molecule? That is, predict one or two key IR absorptions for the reactants that would distinguish it/them from IR absorptions predicted for the product.

Challenge Problems

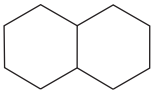
4.47 Consider the cis and trans isomers of 1,3-di-*tert*-butylcyclohexane (build molecular models). What unusual feature accounts for the fact that one of these isomers apparently exists in a twist boat conformation rather than a chair conformation?

4.48 Using the rules found in this chapter, give systematic names for the following or indicate that more rules need to be provided:

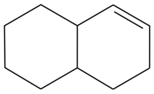
(a)



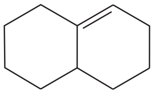
(b)



(c)



(d)



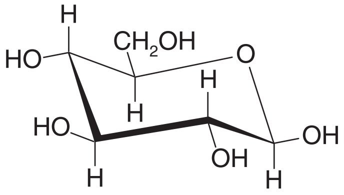
4.49 Open the energy-minimized 3D Molecular Models on the book's website for *trans*-1-*tert*-butyl-3-methylcyclohexane and *trans*-1,3-di-*tert*-butylcyclohexane. What conformations of cyclohexane do the rings in these two compounds resemble most closely? How can you account for the difference in ring conformations between them?

4.50 Open the 3D Molecular Models on the book's website for cyclopentane and vitamin B_{12}. Compare cyclopentane with the nitrogen-containing five-membered rings in vitamin B_{12}. Is the conformation of cyclopentane represented in the specified rings of vitamin B_{12}? What factor(s) account for any differences you observe?

4.51 Open the 3D Molecular Model on the book's website for buckminsterfullerene. What molecule has its type of ring represented 16 times in the surface of buckminsterfullerene?

Learning Group Problems

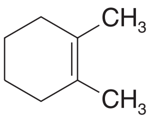
1. This is the predominant conformation for D-glucose:



Why is it not surprising that D-glucose is the most commonly found sugar in nature? (*Hint*: Look up structures for sugars such as D-galactose and D-mannose, and compare these with D-glucose.)

2. Using Newman projections, depict the relative positions of the substituents on the bridgehead atoms of *cis*- and *trans*-decalin. Which of these isomers would be expected to be more stable, and why?

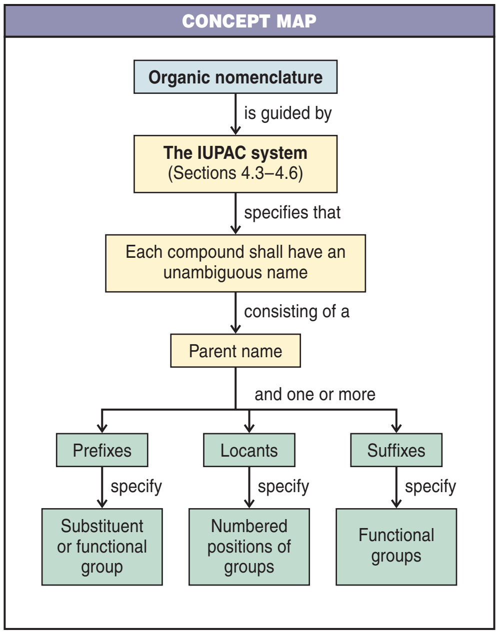
3. When 1,2-dimethylcyclohexene (below) is allowed to react with hydrogen in the presence of a platinum catalyst, the product of the reaction is a cycloalkane that has a melting point of −50°C and a boiling point of 130°C (at 760 torr). **(a)** What is the structure of the product of this reaction? **(b)** Consult an appropriate resource (such as the web or a CRC handbook) and tell which stereoisomer it is. **(c)** What does this experiment suggest about the mode of addition of hydrogen to the double bond?



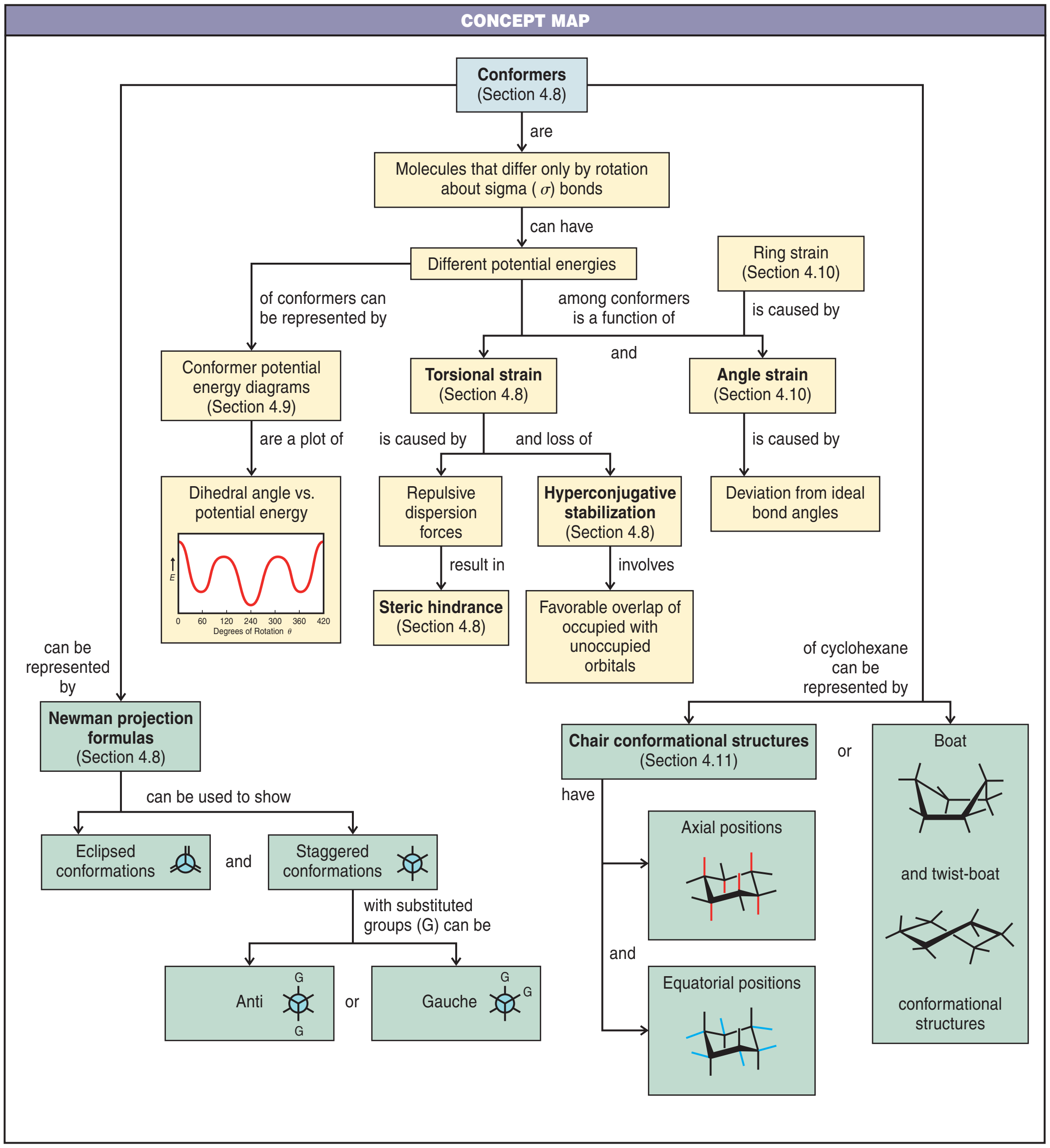
1,2-Dimethylcyclohexene

4. When cyclohexene is dissolved in an appropriate solvent and allowed to react with chlorine, the product of the reaction, $C_6H_{10}Cl_2$, has a melting point of −7°C and a boiling point (at 16 torr) of 74°C. **(a)** Which stereoisomer is this? **(b)** What does this experiment suggest about the mode of addition of chlorine to the double bond?

CONCEPT MAP



CONCEPT MAP

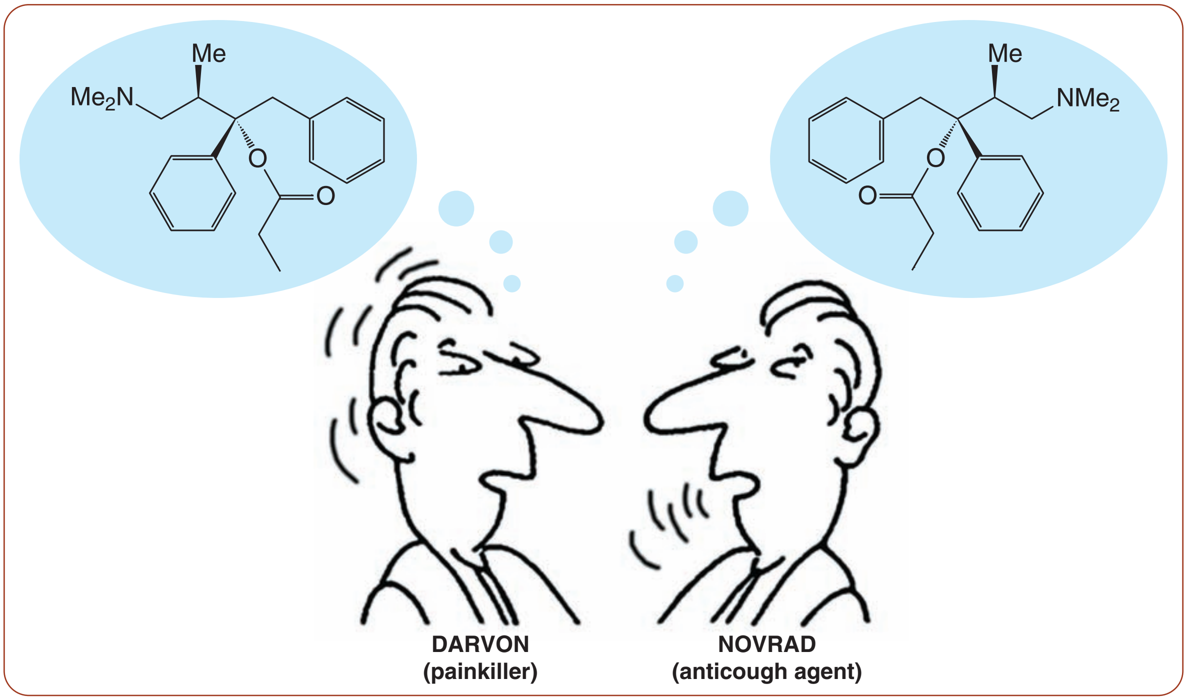


See *Special Topic A* in *WileyPLUS*

5

Stereochemistry

Chiral Molecules



We are all aware of the fact that certain everyday objects such as gloves and shoes possess the quality of "handedness." A right-handed glove only fits a right hand; a left-handed shoe only fits a left foot. Objects that can exist in right-handed and left-handed forms are said to be **chiral**. In this chapter we shall find that molecules can also be chiral and can exist in right- and left-handed forms. For example, one chiral form of the molecule shown above is a painkiller (Darvon), and the other, a cough suppressant (Novrad)! It is easy to see why it is important to understand chirality in molecules.

5.1 *Chirality and Stereochemistry*

The glass and its mirror image are superposable.

Chirality is a phenomenon that pervades the universe. How can we know whether a particular object is **chiral** or **achiral** (not chiral)?

- We can tell if an object has **chirality** by examining the object and its mirror image.

Every object has a mirror image. Many objects are achiral. By this we mean that *the object and its mirror image are identical*, that is, the object and its mirror image are **superposable** one on the other.* Superposable means that one can, in one's mind's eye, place one object on the other so that all parts of each coincide. Simple geometrical objects such as a sphere or a cube are achiral. So is an object like a water glass.

- **A chiral object is one that cannot be superposed on its mirror image.**

---

*To be superposable is different than to be super*im*posable. Any two objects can be superimposed simply by putting one object on top of the other, whether or not the objects are the same. To *superpose* two objects (as in the property of superposition) means, on the other hand, that **all parts of each object must coincide**. The condition of superposability must be met for two things to be **identical**.

Figure 5.1 The mirror image of a right hand is a left hand.

Figure 5.2 Left and right hands are not superposable.

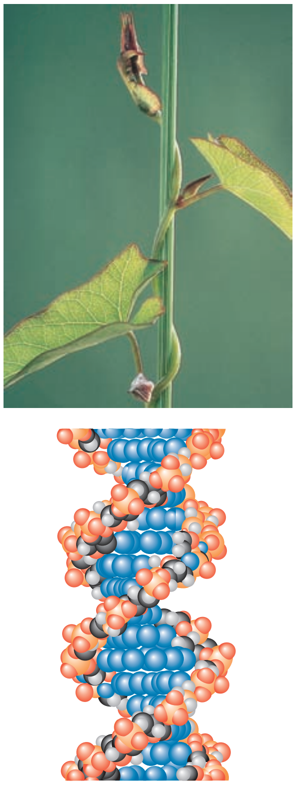
Each of our hands is chiral. When you view your right hand in a mirror, the image that you see in the mirror *is a left hand* (Fig. 5.1). However, as we see in Fig. 5.2, your left hand and your right hand are not identical because *they are not superposable*. Your hands are chiral. In fact, the word chiral comes from the Greek word *cheir* meaning hand. An object such as a mug may or may not be chiral. If it has no markings on it, it is achiral. If the mug has a logo or image on one side, it is chiral.

This mug is chiral.

5.1A The Biological Significance of Chirality

The human body is structurally chiral, with the heart lying to the left of center and the liver to the right. Helical seashells are chiral and most are spiral, such as a right-handed screw. Many plants show chirality in the way they wind around supporting structures. Honeysuckle winds as a left-handed helix; bindweed winds in a right-handed way. DNA is a chiral molecule. The double helical form of DNA turns in a right-handed way.

Chirality in molecules, however, involves more than the fact that some molecules adopt left- or right-handed conformations. As we shall see in this chapter, it is the nature of groups bonded at specific atoms that can bestow chirality on a molecule. Indeed, all but one of the 20 amino acids that make up naturally occurring proteins are chiral, and all of these are classified as being left-handed. The molecules of natural sugars are almost all classified as being right-handed. In fact, most of the molecules of life are chiral, and most are found in only one mirror image form.*



Bindweed (top photo) *(Convolvulus sepium)* winds in a right-handed fashion, like the right-handed helix of DNA.

---

* For interesting reading, see Hegstrum, R. A.; Kondepudi, D. K. The Handedness of the Universe. *Sci. Am.* **1990**, *262*(1), 98–105, and Horgan, J. The Sinister Cosmos. *Sci. Am.* **1997**, *276*(5), 18–19.

Chirality has tremendous importance in our daily lives. Most pharmaceuticals are chiral. Usually only one mirror-image form of a drug provides the desired effect. The other mirror-image form is often inactive or, at best, less active. In some cases the other mirror-image form of a drug actually has severe side effects or toxicity (see Section 5.5 regarding thalidomide). Our senses of taste and smell also depend on chirality. As we shall see, one mirror-image form of a chiral molecule may have a certain odor or taste while its mirror image smells and tastes completely different. The food we eat is largely made of molecules of one mirror-image form. If we were to eat food that was somehow made of molecules with the unnatural mirror-image form, we would likely starve because the enzymes in our bodies are chiral and preferentially react with the natural mirror-image form of their substrates.

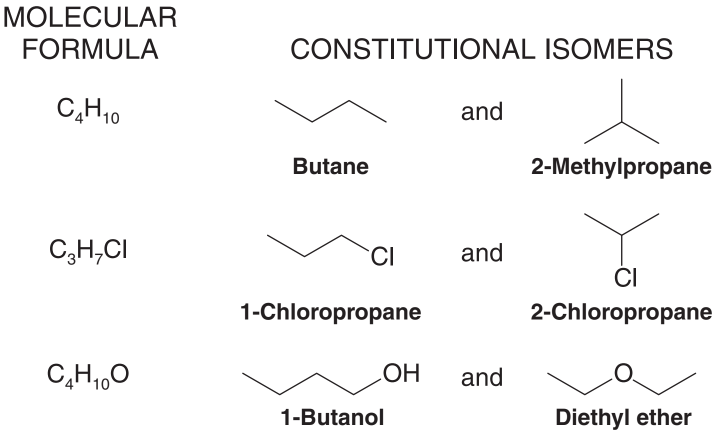
Let us now consider what causes some molecules to be chiral. To begin, we will return to aspects of isomerism.

5.2 *Isomerism: Constitutional Isomers and Stereoisomers*

5.2A Constitutional Isomers

Isomers are different compounds that have the same molecular formula. In our study thus far, much of our attention has been directed toward isomers we have called constitutional isomers.

- **Constitutional isomers** have the same molecular formula but different connectivity, meaning that their atoms are connected in a different order. Examples of constitutional isomers are the following:



5.2B Stereoisomers

Stereoisomers are not constitutional isomers.

- **Stereoisomers** have their atoms connected in the same sequence (the same constitution), but they differ in the arrangement of their atoms in space. The consideration of such spatial aspects of molecular structure is called **stereochemistry**.

We have already seen examples of some types of stereoisomers. The cis and trans forms of alkenes are stereoisomers (Section 1.13B), as are the cis and trans forms of substituted cyclic molecules (Section 4.13).

5.2C Enantiomers and Diastereomers

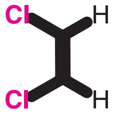
Stereoisomers can be subdivided into two general categories: those that are **enantiomers** of each other, and those that are **diasteromers** of each other.

- **Enantiomers are stereoisomers whose molecules are nonsuperposable mirror images of each other**.

All other stereoisomers are diastereomers.

- **Diastereomers are stereoisomers whose molecules are not mirror images of each other**.

The alkene isomers *cis*- and *trans*-1,2-dichloroethene shown here are stereoisomers that are **diastereomers**.

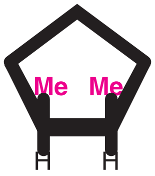


***cis*-1,2-Dichloroethene**
($C_2H_2Cl_2$)

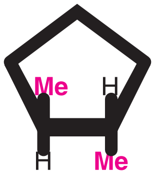
***trans*-1,2-Dichloroethene**
($C_2H_2Cl_2$)

By examining the structural formulas for *cis*- and *trans*-1,2-dichloroethene, we see that they have the same molecular formula ($C_2H_2Cl_2$) and the same connectivity (both compounds have two central carbon atoms joined by a double bond, and both compounds have one chlorine and one hydrogen atom attached to each carbon atom). But, their atoms have a different arrangement in space that is not interconvertible from one to another (due to the large barrier to rotation of the carbon–carbon double bond), making them stereoisomers. Furthermore, they are stereoisomers that are not mirror images of each other; therefore they are diastereomers and not enantiomers.

Cis and trans isomers of cycloalkanes furnish us with another example of stereoisomers that are diastereomers. Consider the following two compounds:



***cis*-1,2-Dimethylcyclopentane**
(C_7H_{14})

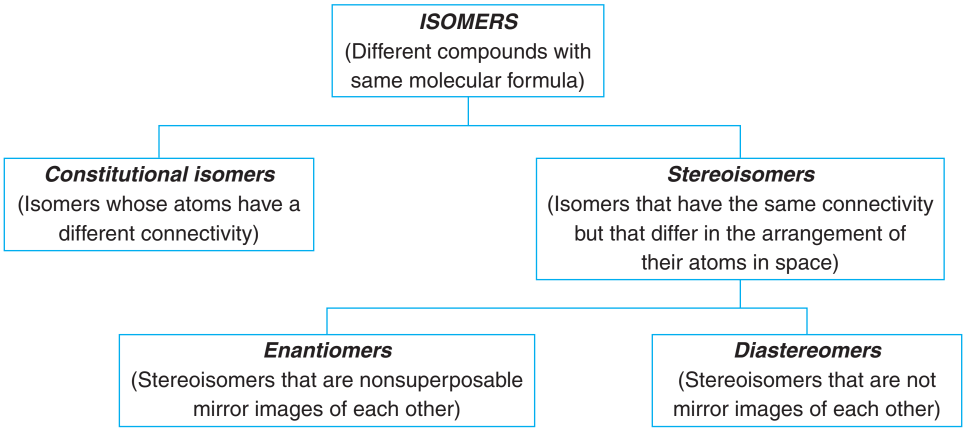


***trans*-1,2-Dimethylcyclopentane**
(C_7H_{14})

These two compounds have the same molecular formula (C_7H_{14}), the same sequence of connections for their atoms, but different arrangements of their atoms in space. In one compound both methyl groups are bonded to the same face of the ring, while in the other compound the two methyl groups are bonded to opposite faces of the ring. Furthermore, the positions of the methyl groups cannot be interconverted by conformational changes. Therefore, these compounds are stereoisomers, and because they are stereoisomers that are not mirror images of each other, they can be further classified as diastereomers.

In Section 5.12 we shall study other molecules that can exist as diastereomers but are not cis and trans isomers of each other. First, however, we need to consider enantiomers further.

SUBDIVISION OF ISOMERS



ISOMERS
(Different compounds with same molecular formula)

Constitutional isomers
(Isomers whose atoms have a different connectivity)

Stereoisomers
(Isomers that have the same connectivity but that differ in the arrangement of their atoms in space)

Enantiomers
(Stereoisomers that are nonsuperposable mirror images of each other)

Diastereomers
(Stereoisomers that are not mirror images of each other)

5.3 *Enantiomers and Chiral Molecules*

Enantiomers always have the possibility of existing in pairs. We may not always find that nature (or a reaction) has produced a pair of enantiomers, however. In fact, in nature we often find only one enantiomer of the two that are possible. We shall find out later why this is often the case. Typically, when we carry out a chemical reaction, we find that the reaction produces a pair of enantiomers. Again, we will explain later why this occurs.

What structural feature must be present for two molecules to exist as enantiomers?

- **Enantiomers occur only with compounds whose molecules are chiral**.

How do we recognize a chiral molecule?

- **A chiral molecule is one that is not superposable on its mirror image**.

What is the relationship between a chiral molecule and its mirror image?

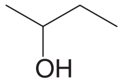
- **The relationship is one that is enantiomeric**. A chiral molecule and its mirror image are said to be enantiomers of each other.

Review Problem 5.1

Classify each of the following objects as to whether it is chiral or achiral:

(a) A screwdriver
(b) A baseball bat
(c) A golf club
(d) A tennis shoe
(e) An ear
(f) A woodscrew
(g) A car
(h) A hammer

The chirality of molecules can be demonstrated with relatively simple compounds. Consider, for example, 2-butanol:



2-Butanol

Until now, we have presented the formula for 2-butanol as though it represented only one compound and we have not mentioned that molecules of 2-butanol are chiral. Because they are, there are actually two different 2-butanols and these two 2-butanols are enantiomers. We can understand this if we examine the drawings and models in Fig. 5.3.

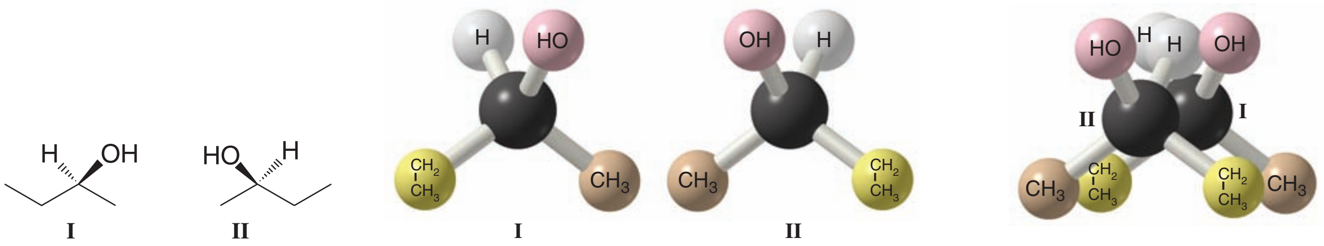
If model **I** is held before a mirror, model **II** is seen in the mirror and vice versa. Models **I** and **II** are not superposable on each other; therefore, they represent different, but isomeric, molecules. ***Because models I and II are nonsuperposable mirror images of each other, the molecules that they represent are enantiomers***.

Helpful Hint

Working with models is a helpful study technique whenever three-dimensional aspects of chemistry are involved.

Review Problem 5.2

Construct handheld models of the 2-butanols represented in Fig. 5.3 and demonstrate for yourself that they are not mutually superposable. **(a)** Make similar models of 2-bromopropane. Are they superposable? **(b)** Is a molecule of 2-bromopropane chiral? **(c)** Would you expect to find enantiomeric forms of 2-bromopropane?



(*a*) (*b*) (*c*)

Figure 5.3 (*a*) Three-dimensional drawings of the 2-butanol enantiomers **I** and **II**. (*b*) Models of the 2-butanol enantiomers. (*c*) An unsuccessful attempt to superpose models of **I** and **II**.

5.4 *A Single Chirality Center Causes a Molecule to Be Chiral*

What structural feature can we use to know when to expect the possible existence of a pair of enantiomers?

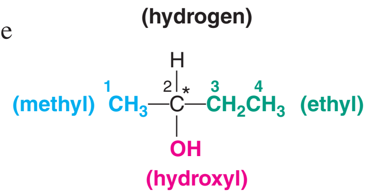
- One way (but not the only way) is to recognize that *a pair of enantiomers is always possible for molecules that contain* **a single tetrahedral atom with four different groups attached to it**.

Traditionally such atoms have been called *asymmetric atoms*, or *stereogenic atoms*, or *stereocenters*. In 1996, however, the IUPAC recommended that such atoms be called **chirality centers**, and this is the usage that we shall follow in this text.* It is also important to state that chirality is a property of the molecule as a whole, and that a chirality center is a structural feature that can cause a molecule to be chiral.

Chirality centers are often designated with an asterisk (*). In 2-butanol the chirality center is C2 (Fig. 5.4). The four different groups that are attached to C2 are a hydroxyl group, a hydrogen atom, a methyl group, and an ethyl group.

An ability to find chirality centers in structural formulas will help us in recognizing molecules that are chiral, and that can exist as enantiomers. **The presence of a single chirality center in a molecule guarantees that the molecule is chiral and that enantiomeric forms are possible**. *However, as we shall see in Section 5.12, there are molecules with more than one chirality center that are not chiral, and there are molecules that do not contain a chirality center that are chiral.*

Figure 5.5 demonstrates that enantiomeric compounds can exist whenever a molecule contains a single chirality center.



$$\overset{1}{CH_3}-\underset{OH}{\overset{H}{\overset{2|\,*}{C}}}-\overset{3}{CH_2}\overset{4}{CH_3}$$

Figure 5.4 The tetrahedral carbon atom of 2-butanol that bears four different groups. [By convention, chirality centers are often designated with an asterisk (*).]

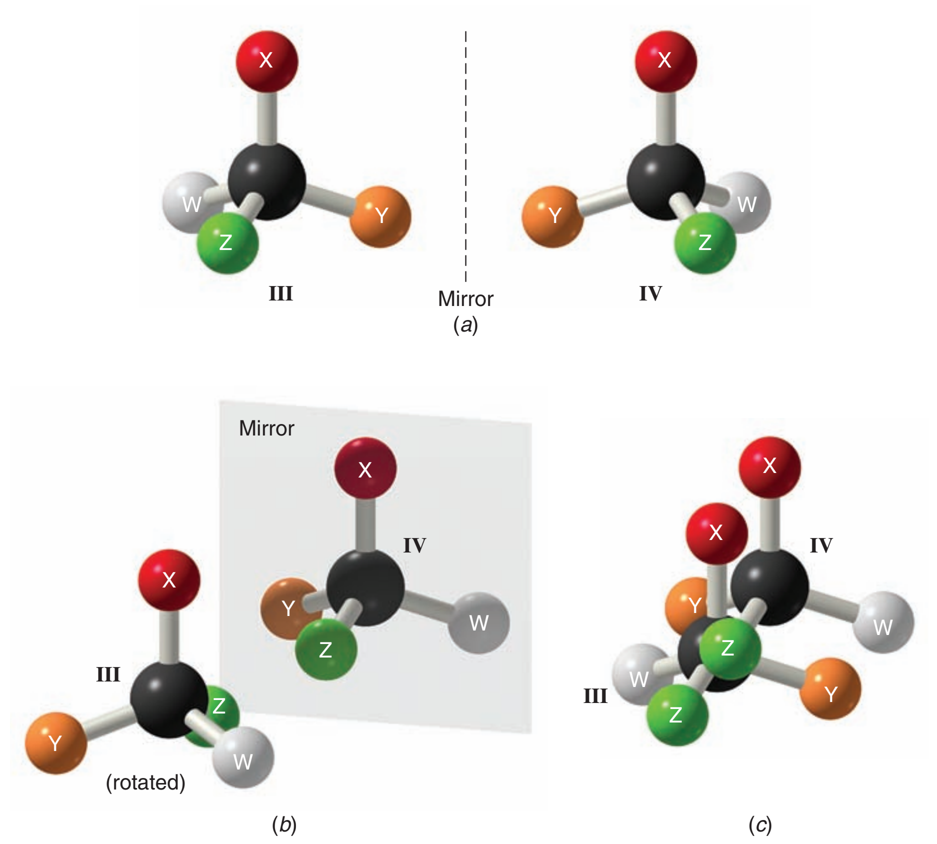


Figure 5.5 A demonstration of chirality of a generalized molecule containing one chirality center. *(a)* The four different groups around the carbon atom in **III** and **IV** are arbitrary. *(b)* **III** is rotated and placed in front of a mirror. **III** and **IV** are found to be related as an object and its mirror image. *(c)* **III** and **IV** are not superposable; therefore, the molecules that they represent are chiral and are enantiomers.

---

*The 1996 IUPAC recommended usage can be found at http://www.chem.qmw.ac.uk/iupac/stereo.

Helpful Hint

Interchanging two groups of a model or three-dimensional formula is a useful test for determining whether structures of two chiral molecules are the same or different.

An important property of enantiomers with a single chirality center, such as 2-butanol, is that *interchanging any two groups at the chirality center converts one enantiomer into the other*. In Fig. 5.3*b* it is easy to see that interchanging the methyl and ethyl groups converts one enantiomer into the other. You should now convince yourself that interchanging any other two groups has the same result.

- Any atom at which an interchange of groups produces a stereoisomer is called a **stereogenic center**. (If the atom is a carbon atom it is usually called a **stereogenic carbon**.)

When we discuss interchanging groups like this, we must take care to notice that what we are describing is *something we do to a molecular model* or *something we do on paper*. An interchange of groups in a real molecule, if it can be done, requires breaking covalent bonds, and this is something that requires a large input of energy. This means that enantiomers such as the 2-butanol enantiomers ***do not interconvert*** spontaneously.

The *chirality center* of 2-butanol is one example of a *stereogenic center*, but there are stereogenic centers that are *not* chirality centers. The carbon atoms of *cis*-1,2-dichloroethene and of *trans*-1,2-dichloroethene (Section 5.2c) are stereogenic centers because an interchange of groups at either carbon atom produces the other stereoisomer. The carbon atoms of *cis*-1,2-dichloroethene and *trans*-1,2-dichloroethene are not chirality centers, however, because they do not have four different groups attached to them.

Review Problem 5.3

Demonstrate the validity of what we have represented in Fig. 5.5 by constructing models. Demonstrate for yourself that **III** and **IV** are related as an object and its mirror image *and that they are not superposable* (i.e., that **III** and **IV** are chiral molecules and are enantiomers). **(a)** Take **IV** and exchange the positions of any two groups. What is the new relationship between the molecules? **(b)** Now take either model and exchange the positions of any two groups. What is the relationship between the molecules now?

If all of the tetrahedral atoms in a molecule have two or more groups attached that *are the same*, the molecule does not have a chirality center. The molecule is superposable on its mirror image and is **achiral**. An example of a molecule of this type is 2-propanol; carbon atoms 1 and 3 bear three identical hydrogen atoms and the central atom bears two identical methyl groups. If we write three-dimensional formulas for 2-propanol, we find (Fig. 5.6) that one structure can be superposed on its mirror image.

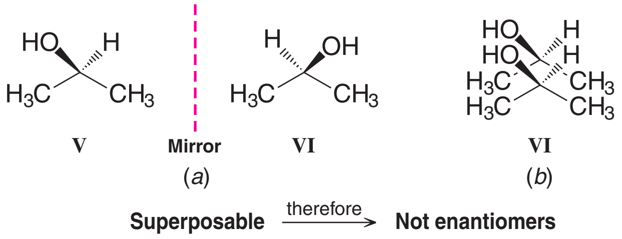


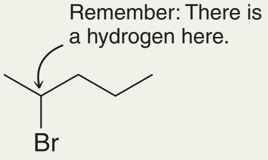
Figure 5.6 *(a)* 2-Propanol (**V**) and its mirror image (**VI**). *(b)* When either one is rotated, the two structures are superposable and so do not represent enantiomers. They represent two molecules of the same compound. 2-Propanol does not have a chirality center.

Thus, we would not predict the existence of enantiomeric forms of 2-propanol, and experimentally only one form of 2-propanol has ever been found.

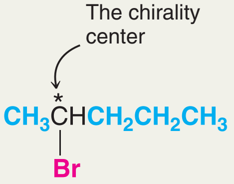
Solved Problem 5.1

Does 2-bromopentane have a chirality center? If so, write three-dimensional structures for each enantiomer.

STRATEGY AND ANSWER First we write a structural formula for the molecule and look for a carbon atom that has four different groups attached to it. In this case, carbon 2 has four different groups: a hydrogen, a methyl group, a bromine, and a propyl group. Thus, carbon 2 is a **chirality center**.



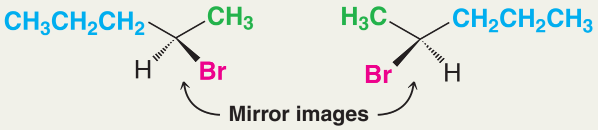
or



$\underset{\text{Br}}{\text{CH}_3\overset{*}{\text{C}}\text{HCH}_2\text{CH}_2\text{CH}_3}$

2-Bromopentane

The enantiomers are



Review Problem 5.4

Some of the molecules listed here have a chirality center; some do not. Write three-dimensional formulas for both enantiomers of those molecules that do have a chirality center.

(a) 2-Fluoropropane
(b) 2-Methylbutane
(c) 2-Chlorobutane
(d) 2-Methyl-1-butanol
(e) *trans*-2-Butene
(f) 2-Bromopentane
(g) 3-Methylpentane
(h) 3-Methylhexane
(i) 2-Methyl-2-pentene
(j) 1-Chloro-2-methylbutane

5.4A Tetrahedral versus Trigonal Stereogenic Centers

It is important to clarify the difference between stereogenic centers, in general, and a chirality center, which is one type of stereogenic center. The chirality center in 2-butanol is a tetrahedral stereogenic center. The carbon atoms of *cis*- and *trans*-1,2-dichloroethene are also stereogenic centers, but they are trigonal stereogenic centers. They are *not* chirality centers. An interchange of groups at the alkene carbons of either 1,2-dichloroethene isomer produces a stereoisomer (a molecule with the same connectivity but a different arrangement of atoms in space), but it does not produce a nonsuperposable mirror image. A chirality center, on the other hand, is one that must have the possibility of nonsuperposable mirror images.

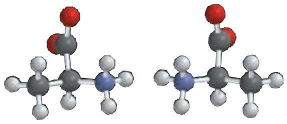
- Chirality centers are tetrahedral stereogenic centers.
- Cis and trans alkene isomers contain trigonal stereogenic centers.

THE CHEMISTRY OF . . .

Life's Molecular Handedness

The amino acids that make up our proteins possess "handedness." They are chiral. Although both mirror image forms are possible, such as those shown below for the amino acid alanine, life on Earth involves amino acids whose chirality is "left-handed" (designated L). The reason that most amino acids are of the left-handed form is not known, however.

In the absence of an influence that possesses handedness such as a living system, chemical reactions produce an equal mixture of both mirror-image forms. Since almost all theories about the origin of life presume that amino acids and other molecules central to life were present before self-replicating organisms came into being, it was assumed that they were present in equal mirror-image forms in the primordial soup.



Alanine enantiomers

But could the mirror-image forms of these molecules actually have been present in unequal amounts before life began, leading to some sort of preference as life evolved? A meteorite discovered in 1970, known as the Murchison meteorite, fueled speculation about this topic. Analysis of the meteorite showed that amino acids and other complex molecules associated with life were present, proving that molecules required for life could arise outside the confines of Earth. Even more interesting, recent experiments have shown that a 7–9% excess of four L-amino acids is present in the Murchison meteorite. The origin of this unequal distribution is uncertain, but some scientists speculate that electromagnetic radiation emitted in a corkscrew fashion from the poles of spinning neutron stars could lead to a bias of one mirror-image isomer over another when molecules form in interstellar space.

5.5 *More about the Biological Importance of Chirality*

The origin of biological properties relating to chirality is often likened to the specificity of our hands for their respective gloves; the binding specificity for a chiral molecule (like a hand) at a chiral receptor site (a glove) is only favorable in one way. If either the molecule or the biological receptor site had the wrong handedness, the natural physiological response (e.g., neural impulse, reaction catalysis) would not occur. A diagram showing how only one amino acid in a pair of enantiomers can interact in an optimal way with a hypothetical binding site (e.g., in an enzyme) is shown in Fig. 5.7. Because of the chirality center of the amino acid, three-point binding can occur with proper alignment for only one of the two enantiomers.

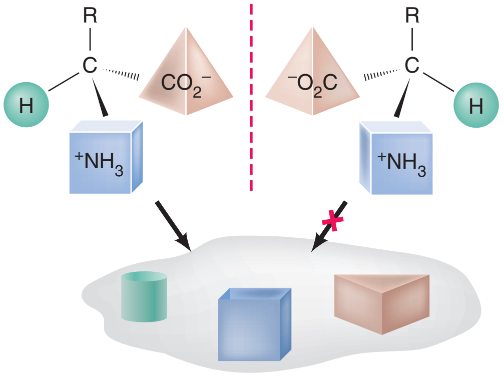
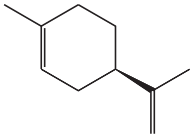
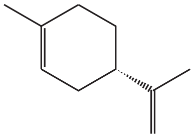


Figure 5.7 Only one of the two amino acid enantiomers shown (the left-hand one) can achieve three-point binding with the hypothetical binding site (e.g., in an enzyme).

Chiral molecules can show their handedness in many ways, including the way they affect human beings. One enantiomeric form of a compound called limonene (Section 23.3) is primarily responsible for the odor of oranges and the other enantiomer for the odor of lemons.



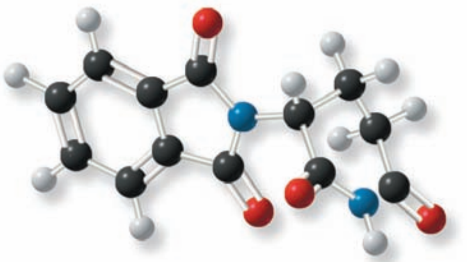
(+)-Limonene (the enantiomer of limonene found in oranges)



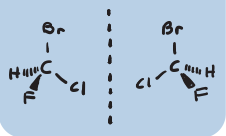
(−)-Limonene (the enantiomer of limonene found in lemons)

One enantiomer of a compound called carvone (Review Problem 5.14) is the essence of caraway, and the other is the essence of spearmint.

The activity of drugs containing chirality centers can similarly vary between enantiomers, sometimes with serious or even tragic consequences. For several years before 1963 the drug thalidomide was used to alleviate the symptoms of morning sickness in pregnant women. In 1963 it was discovered that thalidomide was the cause of horrible birth defects in many children born subsequent to the use of the drug.



Thalidomide (Thalomid®)



Even later, evidence began to appear indicating that whereas one of the thalidomide enantiomers (the right-handed molecule) has the intended effect of curing morning sickness, the other enantiomer, which was also present in the drug (in an equal amount), may be the cause of the birth defects. The evidence regarding the effects of the two enantiomers is complicated by the fact that, under physiological conditions, the two enantiomers are interconverted. Now, however, thalidomide is approved under highly strict regulations for treatment of some forms of cancer and a serious complication associated with leprosy. Its potential for use against other conditions including AIDS and rheumatoid arthritis is also under investigation. We shall consider other aspects of chiral drugs in Section 5.11.

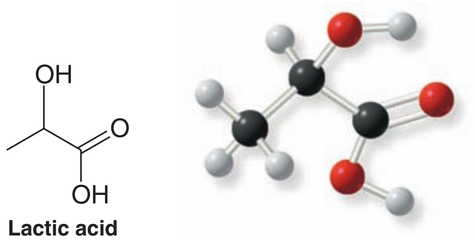
Review Problem 5.5

Which atom is the chirality center of **(a)** limonene and **(b)** of thalidomide?

Review Problem 5.6

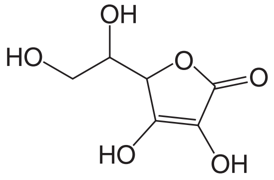
Which atoms in each of the following molecules are chirality centers?

(a)



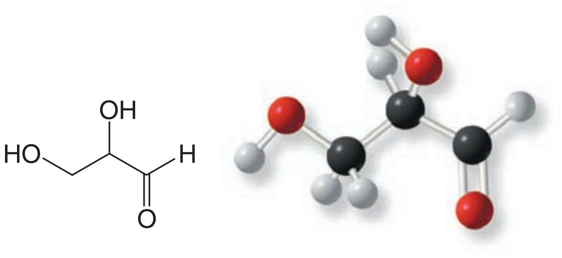
Lactic acid

(c)



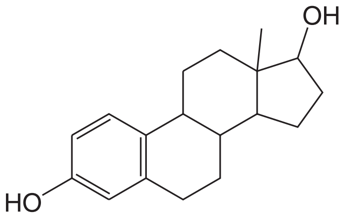
Ascorbic acid (vitamin C)

(b)



Glyceraldehyde

(d)



Estradiol (an estrogen)

5.6 How to Test for Chirality: Planes of Symmetry

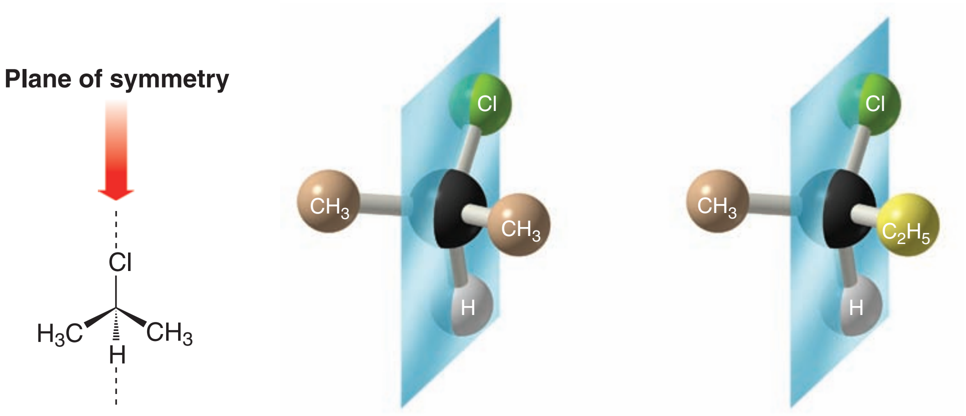
The ultimate way to test for molecular chirality is to construct models of the molecule and its mirror image and then determine whether they are superposable. If the two models are superposable, the molecule that they represent is achiral. If the models are not superposable, then the molecules that they represent are chiral. We can apply this test with actual models, as we have just described, or we can apply it by drawing three-dimensional structures and attempting to superpose them in our minds.

There are other aids, however, that will assist us in recognizing chiral molecules. We have mentioned one already: **the presence of a *single* chirality center**. Other aids are based on the absence of certain symmetry elements in the molecule.

- A molecule will not be chiral if it possesses a plane of symmetry.
- A plane of symmetry (also called a mirror plane) is defined as an imaginary plane that bisects a molecule in such a way that the two halves of the molecule are mirror images of each other.

The plane may pass through atoms, between atoms, or both. For example, 2-chloropropane has a plane of symmetry (Fig. 5.8*a*), whereas 2-chlorobutane does not (Fig. 5.8*b*).

- All molecules with a plane of symmetry in their most symmetric conformation are achiral.



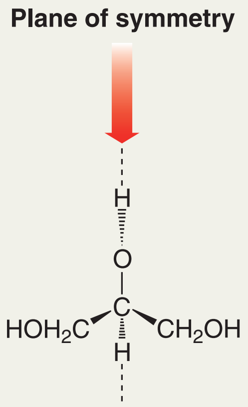
Achiral (*a*)

Chiral (*b*)

Figure 5.8 *(a)* 2-Chloropropane has a plane of symmetry and is achiral. *(b)* 2-Chlorobutane does not possess a plane of symmetry and is chiral.

Solved Problem 5.2

Glycerol, $CH_2OHCHOHCH_2OH$, is an important constituent in the biological synthesis of fats, as we shall see in Chapter 23. (**a**) Does glycerol have a plane of symmetry? If so, write a three-dimensional structure for glycerol and indicate where it is. (**b**) Is glycerol chiral?



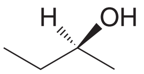
STRATEGY AND ANSWER (**a**) Yes, glycerol has a plane symmetry. Notice we have to choose the proper conformation and orientation of the molecule to see the plane of symmetry. (**b**) No, it is achiral because it has a conformation containing a plane of symmetry.

Review Problem 5.7 Which of the objects listed in Review Problem 5.1 possess a plane of symmetry and are, therefore, achiral?

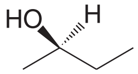
Review Problem 5.8 Write three-dimensional formulas and designate a plane of symmetry for all of the achiral molecules in Review Problem 5.4. (In order to be able to designate a plane of symmetry you may need to write the molecule in an appropriate conformation.

5.7 *Naming Enantiomers: The R,S-System*

The two enantiomers of 2-butanol are the following:

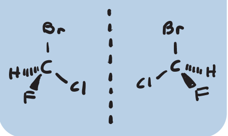


I



II

If we name these two enantiomers using only the IUPAC system of nomenclature that we have learned so far, both enantiomers will have the same name: 2-butanol (or *sec*-butyl alcohol) (Section 4.3F). This is undesirable because *each compound must have its own distinct name*. Moreover, the name that is given a compound should allow a chemist



who is familiar with the rules of nomenclature to write the structure of the compound from its name alone. Given the name 2-butanol, a chemist could write either structure **I** or structure **II**.

Three chemists, R. S. Cahn (England), C. K. Ingold (England), and V. Prelog (Switzerland), devised a system of nomenclature that, when added to the IUPAC system, solves both of these problems. This system, called the ***R*,*S*-system** or the Cahn–Ingold–Prelog system, is part of the IUPAC rules.

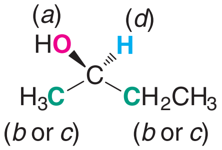
According to this system, one enantiomer of 2-butanol should be designated (*R*)-2-butanol and the other enantiomer should be designated (*S*)-2-butanol. [(*R*) and (*S*) are from the Latin words *rectus* and *sinister*, meaning right and left, respectively.] These molecules are said to have opposite **configurations** at C2.

5.7A How to Assign (*R*) and (*S*) Configurations

We assign (*R*) and (*S*) configurations on the basis of the following procedure.

1. Each of the four groups attached to the chirality center is assigned a **priority** or **preference** *a*, *b*, *c*, or *d*. Priority is first assigned on the basis of the **atomic number** of the atom that is directly attached to the chirality center. The group with the lowest atomic number is given the lowest priority, *d*; the group with next higher atomic number is given the next higher priority, *c*; and so on. (In the case of isotopes, the isotope of greatest atomic mass has highest priority.)

We can illustrate the application of the rule with the 2-butanol enantiomer, **II**:

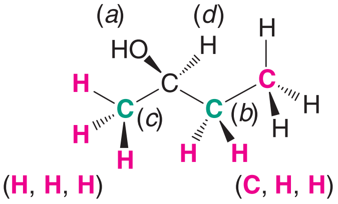


II

Oxygen has the highest atomic number of the four atoms attached to the chirality center and is assigned the highest priority, *a*. Hydrogen has the lowest atomic number and is assigned the lowest priority, *d*. A priority cannot be assigned for the methyl group and the ethyl group by this approach because the atom that is directly attached to the chirality center is a carbon atom in both groups.

2. When a priority cannot be assigned on the basis of the atomic number of the atoms that are directly attached to the chirality center, then the next set of atoms in the unassigned groups is examined. This process is continued until a decision can be made. *We assign a priority at the first point of difference.**

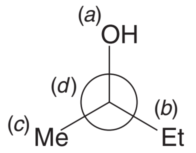
When we examine the methyl group of enantiomer **II**, we find that the next set of atoms consists of three hydrogen atoms (**H, H, H**). In the ethyl group of **II** the next set of atoms consists of one carbon atom and two hydrogen atoms (**C, H, H**). Carbon has a higher atomic number than hydrogen, so we assign the ethyl group the higher priority, *b*, and the methyl group the lower priority, *c*, since (**C, H, H**) > (**H, H, H**):



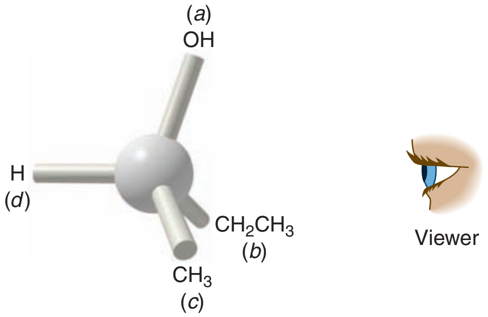
---

*The rules for a branched chain require that we follow the chain with the highest priority atoms.

3. We now rotate the formula (or model) so that the group with lowest priority (*d*) is directed away from us:



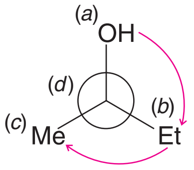
Newman projection



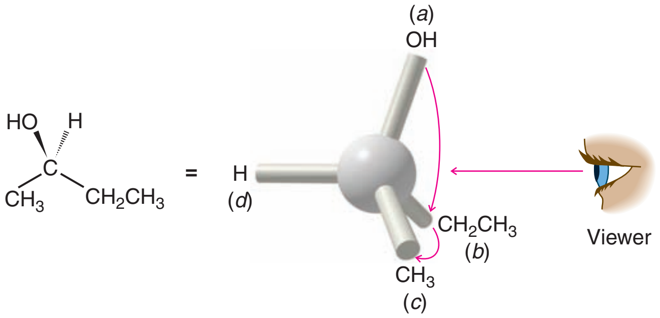
II

Then we trace a path from *a* to *b* to *c*. If, as we do this, the direction of our finger (or pencil) is *clockwise*, the enantiomer is designated (*R*). If the direction is *counterclockwise*, the enantiomer is designated (*S*).

On this basis the 2-butanol enantiomer **II** is (*R*)-2-butanol:



Newman projection

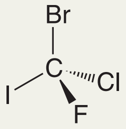


Arrows are clockwise.

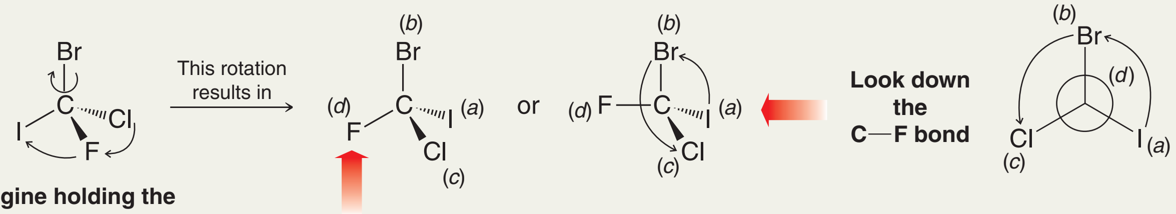
(*R*)-2-Butanol

Solved Problem 5.3

Shown here is an enantiomer of bromochlorofluoroiodomethane. Is it (*R*) or (*S*)?



STRATEGY AND ANSWER



Imagine holding the molecule by the Br and rotating it as shown so that the lowest priority group (F, in this case) lies in the plane of the paper.

Lowest priority

The path we trace from highest to lowest is counterclockwise, so the enantiomer is (*S*).

Review Problem 5.9

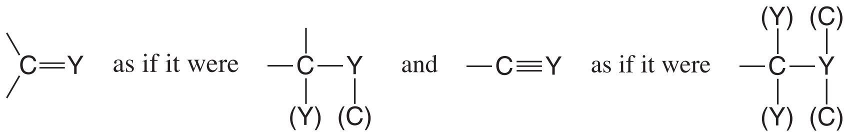
Write the enantiomeric forms of bromochlorofluoromethane and assign each enantiomer its correct (*R*) or (*S*) designation.

Review Problem 5.10

Give (*R*) and (*S*) designations for each pair of enantiomers given as answers to Review Problem 5.4.

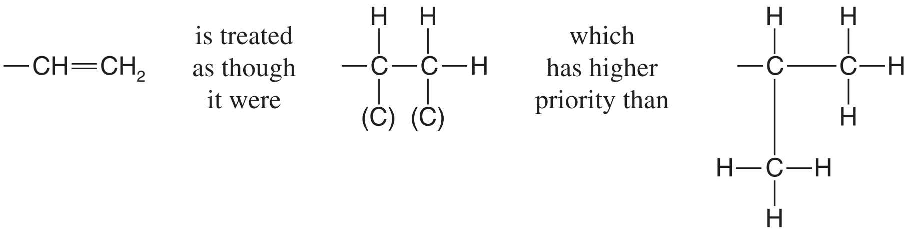
The first three rules of the Cahn–Ingold–Prelog system allow us to make an (*R*) or (*S*) designation for most compounds containing single bonds. For compounds containing multiple bonds one other rule is necessary:

4. Groups containing double or triple bonds are assigned priorities as if both atoms were duplicated or triplicated, that is,

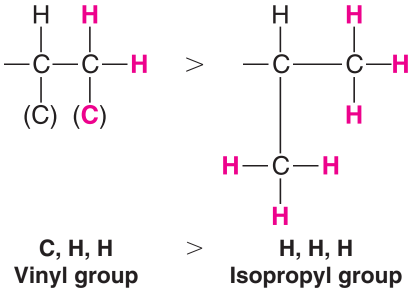


where the symbols in parentheses are duplicate or triplicate representations of the atoms at the other end of the multiple bond.

Thus, the vinyl group, $—CH{=}CH_2$, is of higher priority than the isopropyl group, $—CH(CH_3)_2$. That is,



because at the second set of atoms out, the vinyl group (see the following structure) is **C**, **H**, **H**, whereas the isopropyl group along either branch is **H**, **H**, **H**. (At the first set of atoms both groups are the same: **C**, **C**, **H**.)



Other rules exist for more complicated structures, but we shall not study them here.*

Review Problem 5.11

List the substituents in each of the following sets in order of priority, from highest to lowest:

(a) $—Cl$, $—OH$, $—SH$, $—H$

(b) $—CH_3$, $—CH_2Br$, $—CH_2Cl$, $—CH_2OH$

(c) $—H$, $—OH$, $—CHO$, $—CH_3$

(d) $—CH(CH_3)_2$, $—C(CH_3)_3$, $—H$, $—CH{=}CH_2$

(e) $—H$, $—N(CH_3)_2$, $—OCH_3$, $—CH_3$

(f) $—OH$, $—OPO_3H_2$, $—H$, $—CHO$

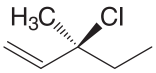
---

*Further information can be found in the Chemical Abstracts Service *Index Guide*.

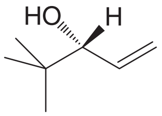
Review Problem 5.12

Assign (*R*) or (*S*) designations to each of the following compounds:

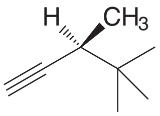
(a)



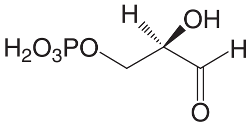
(b)



(c)



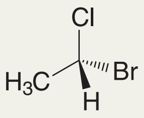
(d)



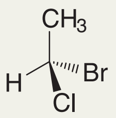
D-Glyceraldehyde-3-phosphate
(a glycolysis intermediate)

Solved Problem 5.4

Consider the following pair of structures and tell whether they represent enantiomers or two molecules of the same compound in different orientations:

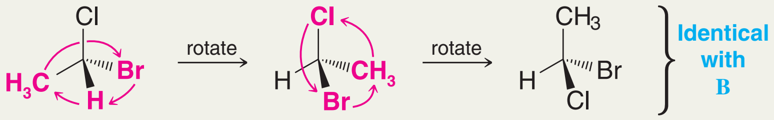


A



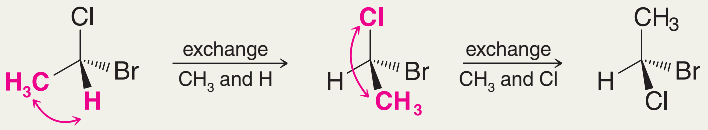
B

STRATEGY One way to approach this kind of problem is to take one structure and, in your mind, hold it by one group. Then rotate the other groups until at least one group is in the same place as it is in the other structure. (Until you can do this easily in your mind, practice with models.) By a series of rotations like this you will be able to convert the structure you are manipulating into one that is either identical with or the mirror image of the other. For example, take **A**, hold it by the Cl atom and then rotate the other groups about the C*—Cl bond until the hydrogen occupies the same position as in **B**. Then hold it by the H and rotate the other groups about the C*—H bond. This will make **B** identical with **A**:



A **A** **A**

Another approach is to recognize that exchanging two groups at the chirality center *inverts the configuration of* that carbon atom and converts a structure *with only one chirality center* into its enantiomer; a second exchange recreates the original molecule. So we proceed this way, keeping track of how many exchanges are required to convert **A** into **B**. In this instance we find that two exchanges are required, and, again, we conclude that **A** and **B** are the same:



A **B**

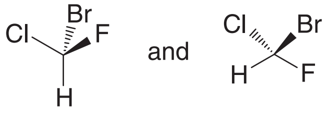
A useful check is to name each compound including its (*R,S*) designation. If the names are the same, then the structures are the same. In this instance both structures are (*R*)-1-bromo-1-chloroethane.

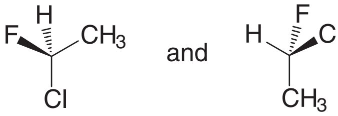
Another method for assigning (*R*) and (*S*) configurations using one's hands as chiral templates has been described (Huheey, J. E. *J. Chem. Educ*. **1986**, *63*, 598–600). Groups at a chirality center are correlated from lowest to highest priority with one's wrist, thumb, index finger, and second finger, respectively. With the ring and little finger closed against the palm and viewing one's hand with the wrist away, if the correlation between the chirality center is with the left hand, the configuration is (*S*), and if with the right hand, (*R*).

ANSWER **A** and **B** are two molecules of the same compound oriented differently.

Review Problem 5.13

Tell whether the two structures in each pair represent enantiomers or two molecules of the same compound in different orientations.

(a)


(b)


(c)


5.8 *Properties of Enantiomers: Optical Activity*

The molecules of enantiomers are not superposable and, on this basis alone, we have concluded that enantiomers are different compounds. How are they different? Do enantiomers resemble constitutional isomers and diastereomers in having different melting and boiling points? The answer is *no*. Pure enantiomers have *identical* melting and boiling points. Do pure enantiomers have different indexes of refraction, different solubilities in common solvents, different infrared spectra, and different rates of reaction with achiral reagents? The answer to each of these questions is also *no*.

Many of these properties (e.g., boiling points, melting points, and solubilities) are dependent on the magnitude of the intermolecular forces operating between the molecules (Section 2.13), and for molecules that are mirror images of each other these forces will be identical. We can see an example of this if we examine Table 5.1, where boiling points of the 2-butanol enantiomers are listed.

TABLE 5.1 Physical Properties of 2-Butanol and Tartaric Acid Enantiomers

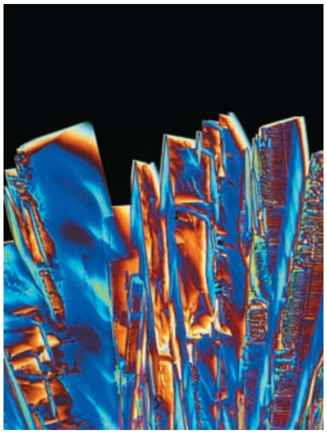
Compound	Boiling Point (bp) or Melting Point (mp)
(*R*)-2-Butanol	99.5°C (bp)
(*S*)-2-Butanol	99.5°C (bp)
(+)-(*R*,*R*)-Tartaric acid	168–170°C (mp)
(−)-(*S*,*S*)-Tartaric acid	168–170°C (mp)
(+/−)-Tartaric acid	210–212°C (mp)

Mixtures of the enantiomers of a compound have different properties than pure samples of each, however. The data in Table 5.1 illustrate this for tartaric acid. The natural isomer, (+)-tartaric acid, has a melting point of 168–170°C, as does its unnatural enantiomer, (−)-tartaric acid. An equal mixture tartaric acid enantiomers, (+/−)-tartaric acid, has a melting point of 210–212°C, however.

Enantiomers show different behavior only when they interact with other chiral substances, including their own enantiomer, as shown by the melting point data above. Enantiomers show different rates of reaction toward other chiral molecules, that is, toward reagents that consist of a single enantiomer or an excess of a single enantiomer. Enantiomers also show different solubilities in solvents that consist of a single enantiomer or an excess of a single enantiomer.

One easily observable way in which enantiomers differ is in *their behavior toward plane-polarized light*. When a beam of plane-polarized light passes through an enantiomer, the plane of polarization **rotates.** Moreover, separate enantiomers rotate the plane of plane-polarized light equal amounts *but in opposite directions*. Because of their effect on plane-polarized light, separate enantiomers are said to be **optically active compounds**.

In order to understand this behavior of enantiomers, we need to understand the nature of plane-polarized light. We also need to understand how an instrument called a polarimeter operates.



Tartaric acid is found naturally in grapes and many other plants. Crystals of tartaric acid can be sometimes be found with wine.

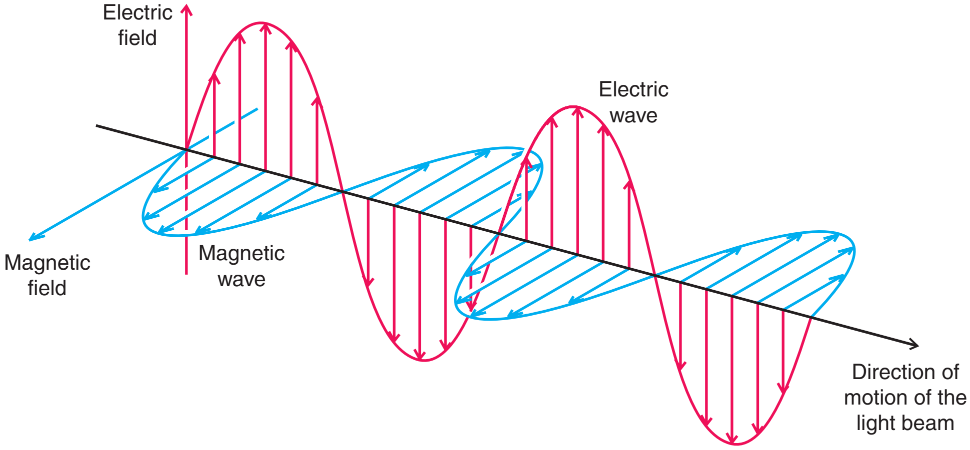


Figure 5.9 The oscillating electric and magnetic fields of a beam of ordinary light in one plane. The waves depicted here occur in all possible planes in ordinary light.

5.8A Plane-Polarized Light

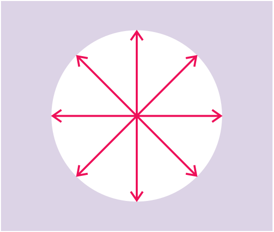


Figure 5.10 Oscillation of the electric field of ordinary light occurs in all possible planes perpendicular to the direction of propagation.

Light is an electromagnetic phenomenon. A beam of light consists of two mutually perpendicular oscillating fields: an oscillating electric field and an oscillating magnetic field (Fig. 5.9).

If we were to view a beam of ordinary light from one end, and if we could actually see the planes in which the electrical oscillations were occurring, we would find that oscillations of the electric field were occurring in all possible planes perpendicular to the direction of propagation (Fig. 5.10). (The same would be true of the magnetic field.)

When ordinary light is passed through a polarizer, the polarizer interacts with the electric field so that the electric field of the light that emerges from the polarizer (and the magnetic field perpendicular to it) is oscillating only in one plane. Such light is called **plane-polarized light** (Fig. 5.11*a*). If the plane-polarized beam encounters a filter with perpendicular polarization, the light is blocked (Fig. 5.11*b*). This phenomenon can readily be demonstrated with lenses from a pair of polarizing sunglasses or a sheet of polarizing film (Fig. 5.11*c*).

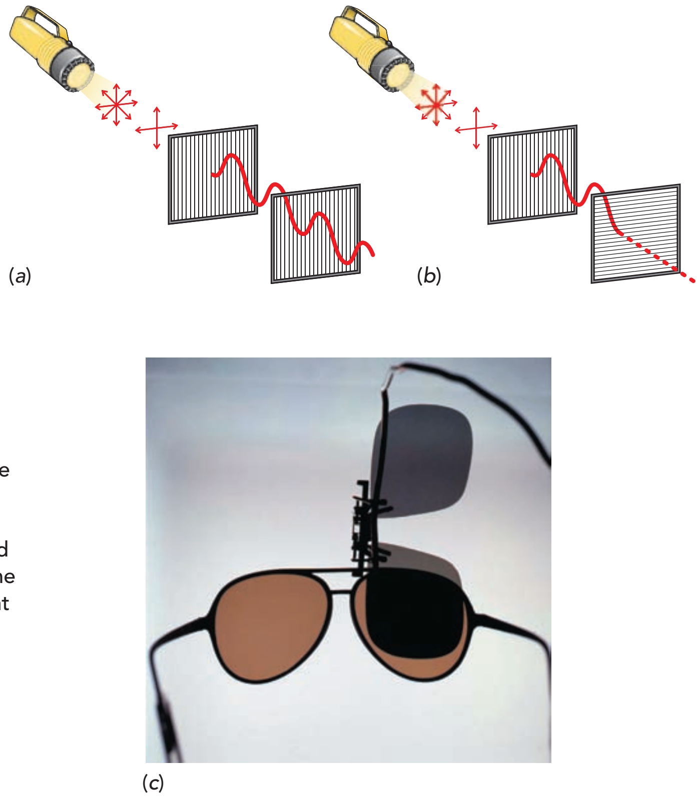
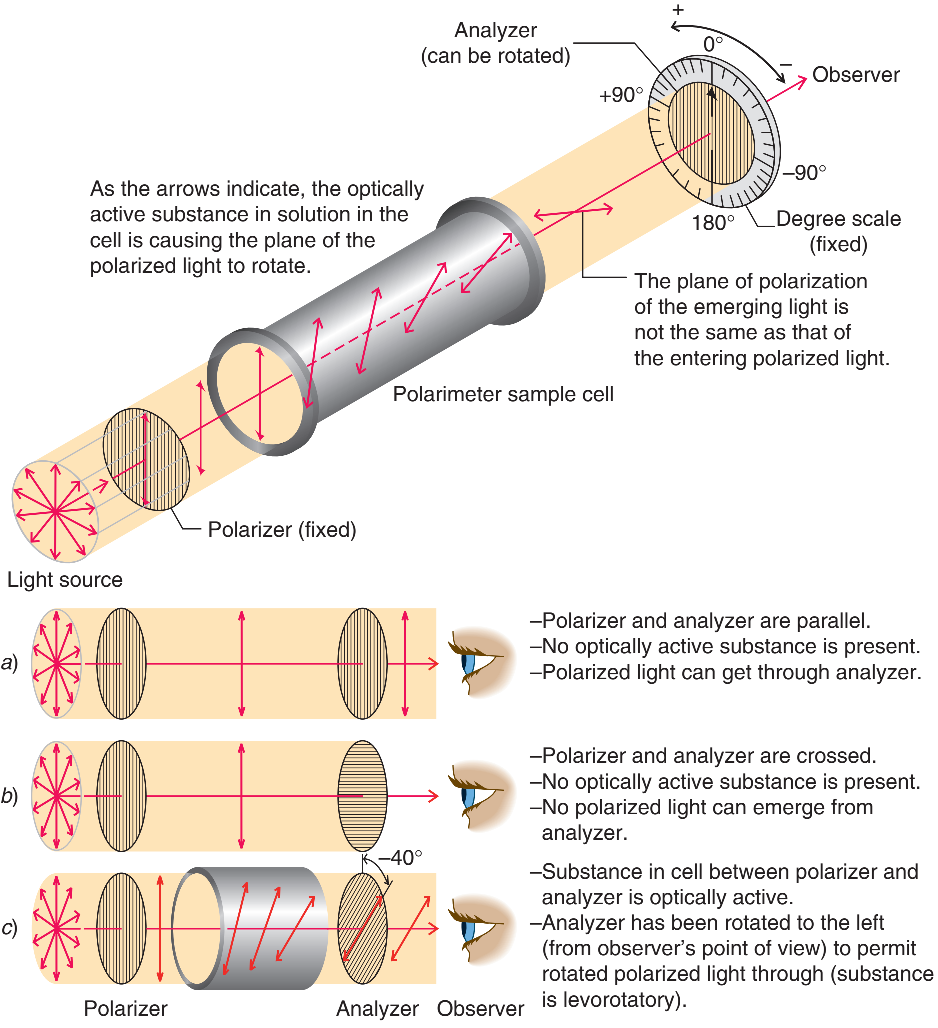


Figure 5.11 *(a)* Ordinary light passing through the first polarizing filter emerges with an electric wave oscillating in only one plane (and a perpendicular magnetic wave plane not shown). When the second filter is aligned with its polarizing direction the same as the first filter, as shown, the plane-polarized light can pass through. *(b)* If the second filter is turned 90°, the plane-polarized light is blocked. *(c)* Two polarizing sunglass lenses oriented perpendicular to each other block the light beam.

5.8B The Polarimeter

The device that is used for measuring the effect of optically active compounds on plane-polarized light is a **polarimeter**. A sketch of a polarimeter is shown in Fig. 5.12. The principal working parts of a polarimeter are (1) a light source (usually a sodium lamp), (2) a polarizer, (3) a cell for holding the optically active substance (or solution) in the light beam, (4) an analyzer, and (5) a scale for measuring the angle (in degrees) that the plane of polarized light has been rotated.

The analyzer of a polarimeter (Fig. 5.12) is nothing more than another polarizer. If the cell of the polarimeter is empty or if an optically *inactive* substance is present, the axes of the plane-polarized light and the analyzer will be exactly parallel when the instrument reads 0°, and the observer will detect the maximum amount of light passing through. If, by contrast, the cell contains an optically active substance, a solution of one enantiomer, for example, the plane of polarization of the light will be rotated as it passes through the cell. In order to detect the maximum brightness of light, the observer will have to rotate the axis of the analyzer in either a clockwise or counterclockwise direction. If the analyzer is rotated in a clockwise direction, the rotation, α (measured in degrees), is said to be positive (+). If the rotation is counterclockwise, the rotation is said to be negative (−). A substance that rotates plane-polarized light in the clockwise direction is also said to be **dextrorotatory**, and one that rotates plane-polarized light in a counterclockwise direction is said to be **levorotatory** (Latin: *dexter*, right, and *laevus*, left).



As the arrows indicate, the optically active substance in solution in the cell is causing the plane of the polarized light to rotate.

The plane of polarization of the emerging light is not the same as that of the entering polarized light.

–Polarizer and analyzer are parallel.
–No optically active substance is present.
–Polarized light can get through analyzer.

–Polarizer and analyzer are crossed.
–No optically active substance is present.
–No polarized light can emerge from analyzer.

–Substance in cell between polarizer and analyzer is optically active.
–Analyzer has been rotated to the left (from observer's point of view) to permit rotated polarized light through (substance is levorotatory).

Figure 5.12 The principal working parts of a polarimeter and the measurement of optical rotation. (Reprinted with permission of John Wiley & Sons, Inc. from Holum, J. R., *Organic Chemistry: A Brief Course*, p. 316. Copyright 1975.)

5.8C Specific Rotation

The number of degrees that the plane of polarization is rotated as the light passes through a solution of an enantiomer depends on the number of chiral molecules that it encounters. This, of course, depends on the length of the tube and the concentration of the enantiomer. In order to place measured rotations on a standard basis, chemists calculate a quantity called the **specific rotation**, $[\alpha]$, by the following equation:

$$[\alpha] = \frac{\alpha}{c \cdot l}$$

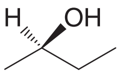
where $[\alpha]$ = the specific rotation
α = the observed rotation
c = the concentration of the solution in grams per milliliter of solution (or density in g mL^{-1} for neat liquids)
l = the length of the cell in decimeters (1 dm = 10 cm)

The specific rotation also depends on the temperature and the wavelength of light that is employed. Specific rotations are reported so as to incorporate these quantities as well. A specific rotation might be given as follows:

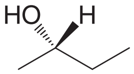
$$[\alpha]_{\mathrm{D}}^{25} = +3.12$$

This means that the D line of a sodium lamp (λ = 589.6 nm) was used for the light, that a temperature of 25°C was maintained, and that a sample containing 1.00 g mL^{-1} of the optically active substance, in a 1-dm tube, produced a rotation of 3.12° in a clockwise direction.*

The specific rotations of (*R*)-2-butanol and (*S*)-2-butanol are given here:

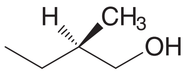


(*R*)-2-Butanol
$[\alpha]_{\mathrm{D}}^{25} = -13.52$

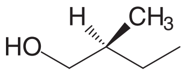


(*S*)-2-Butanol
$[\alpha]_{\mathrm{D}}^{25} = +13.52$

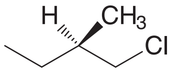
The direction of rotation of plane-polarized light is often incorporated into the names of optically active compounds. The following two sets of enantiomers show how this is done:



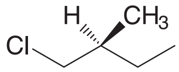
(*R*)-(+)-2-Methyl-1-butanol
$[\alpha]_{\mathrm{D}}^{25} = +5.756$



(*S*)-(−)-2-Methyl-1-butanol
$[\alpha]_{\mathrm{D}}^{25} = -5.756$



(*R*)-(−)-1-Chloro-2-methylbutane
$[\alpha]_{\mathrm{D}}^{25} = -1.64$

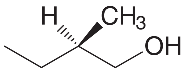


(*S*)-(+)-1-Chloro-2-methylbutane
$[\alpha]_{\mathrm{D}}^{25} = +1.64$

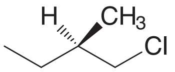
The previous compounds also illustrate an important principle:

- No obvious correlation exists between the (*R*) and (*S*) configurations of enantiomers and the direction [(+) or (−)] in which they rotate plane-polarized light.

(*R*)-(+)-2-Methyl-1-butanol and (*R*)-(−)-1-chloro-2-methylbutane have the same *configuration*; that is, they have the same general arrangement of their atoms in space. They have, however, an opposite effect on the direction of rotation of the plane of plane-polarized light:



Same configuration



(*R*)-(+)-2-Methyl-1-butanol **(*R*)-(−)-1-Chloro-2-methylbutane**

---

*The magnitude of rotation is dependent on the solvent used when solutions are measured. This is the reason the solvent is specified when a rotation is reported in the chemical literature.

These same compounds also illustrate a second important principle:

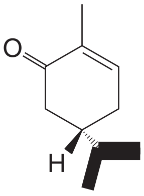
- No necessary correlation exists between the (*R*) and (*S*) designation and the direction of rotation of plane-polarized light.

(*R*)-2-Methyl-1-butanol is dextrorotatory (+), and (*R*)-1-chloro-2-methylbutane is levorotatory (−).

A method based on the measurement of optical rotation at many different wavelengths, called optical rotatory dispersion, has been used to correlate configurations of chiral molecules. A discussion of the technique of optical rotatory dispersion, however, is beyond the scope of this text.

Review Problem 5.14

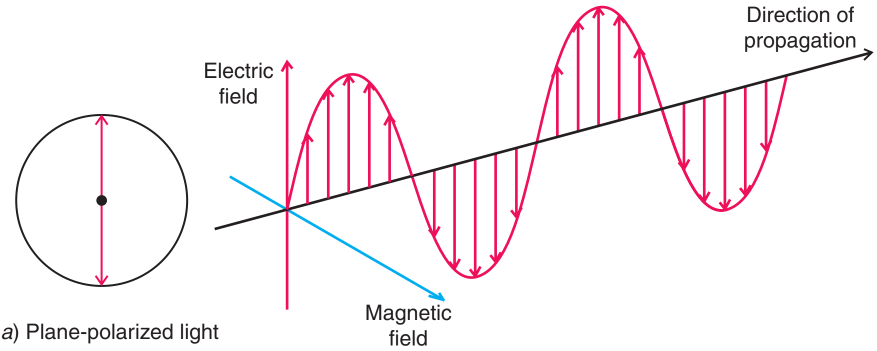
Shown below is the configuration of (+)-carvone. (+)-Carvone is the principal component of caraway seed oil and is responsible for its characteristic odor. (−)-Carvone, its enantiomer, is the main component of spearmint oil and gives it its characteristic odor. The fact that the carvone enantiomers do not smell the same suggests that the receptor sites in the nose for these compounds are chiral, and that only the correct enantiomer binds well to its particular site (just as a hand requires a glove of the correct chirality for a proper fit). Give the correct (*R*) and (*S*) designations for (+)- and (−)-carvone.



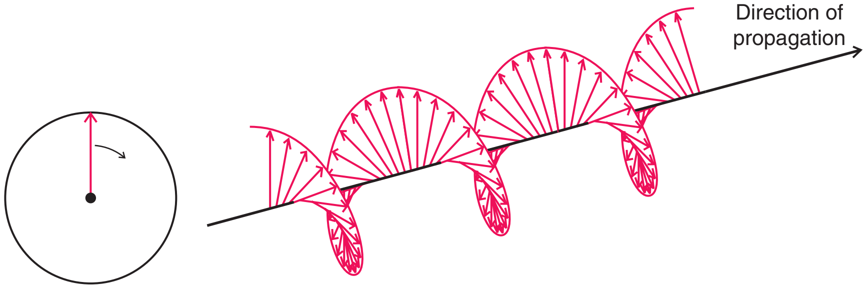
(+)-Carvone

5.9 *The Origin of Optical Activity*

Optical activity is measured by the degree of rotation of plane-polarized light passing through a chiral medium. The theoretical explanation for the origin of optical activity requires consideration of *circularly*-polarized light, however, and its interaction with chiral molecules. While it is not possible to provide a full theoretical explanation for the origin of optical activity here, the following explanation will suffice. A beam of plane-polarized light (Fig. 5.13*a*)



(*a*) Plane-polarized light



(*b*) Circularly-polarized light

Figure 5.13 *(a)* Plane-polarized light. *(b)* Circularly-polarized light. *(c, next page)* Two circularly-polarized beams counterrotating at the same velocity (in phase) and their vector sum. The net result is like *(a)*. *(d, next page)* Two circularly-polarized beams counterrotating at different velocities, such as after interaction with a chiral molecule, and their vector sum. The net result is like *(b)*. Parts *c* and *d:* From ADAMSON. A TEXTBOOK OF PHYSICAL CHEMISTRY, 3E. © 1986 Brooks/Cole, a part of Cengage Learning, Inc. Reproduced by permission. *www.cengage.com/permissions*

(*c*) Two circularly-polarized beams counter-rotating at the same velocity (in phase), and their vector sum. The net result is like (*a*) on the previous page.

(*d*) Two circularly-polarized beams counter-rotating at different velocities, such as after interaction with a chiral molecule, and their vector sum. The net result is like (*b*) on the previous page.

Figure 5.13 *(continued)*

can be described in terms of circularly-polarized light. A beam of circularly-polarized light rotating in one direction is shown in Fig. 5.13*b*. The vector sum of *two* counterrotating in-phase circularly-polarized beams is a beam of plane-polarized light (Fig. 5.13*c*). The optical activity of chiral molecules results from the fact that the two *counterrotating circularly-polarized beams travel with different velocities through the chiral medium*. As the difference between the two circularly-polarized beams propagates through the sample, their vector sum describes a plane that is progressively rotated (Fig. 5.13*d*). What we measure when light emerges from the sample is the net rotation of the plane-polarized light caused by differences in velocity of the circularly-polarized beam components. The origin of the differing velocities has ultimately to do with interactions between electrons in the chiral molecule and light.

Molecules that are not chiral cause no difference in velocity of the two circularly-polarized beams; hence there is no rotation of the plane of polarized light described by their vector sum. Achiral molecules, therefore, are not optically active.

5.9A Racemic Forms

A sample that consists exclusively or predominantly of one enantiomer causes a net rotation of plane-polarized light. Figure 5.14*a* depicts a plane of polarized light as it encounters a molecule of (*R*)-2-butanol, causing the plane of polarization to rotate slightly in one direction. (For the remaining purposes of our discussion we shall limit our description of polarized light to the resultant plane, neglecting consideration of the circularly-polarized components from which plane-polarized light arises.) Each additional molecule of (*R*)-2-butanol that the beam encounters would cause further rotation in the same direction. If, on the other hand, the mixture contained molecules of (*S*)-2-butanol, each molecule of that enantiomer would cause the plane of polarization to rotate in the opposite direction (Fig. 5.14*b*). If the (*R*) and (*S*) enantiomers were present in equal amounts, there would be no net rotation of the plane of polarized light.

- An equimolar mixture of two enantiomers is called a **racemic mixture** (or **racemate** or **racemic form**). **A racemic mixture causes no net rotation of plane-polarized light**.

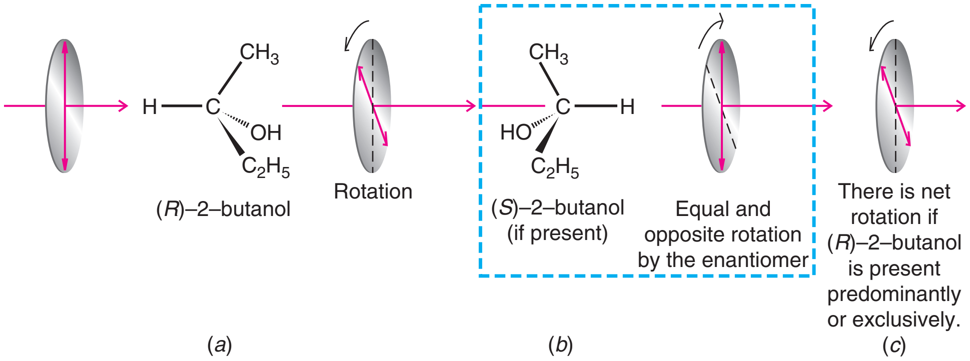


Figure 5.14 *(a)* A beam of plane-polarized light encounters a molecule of (*R*)-2-butanol, a chiral molecule. This encounter produces a slight rotation of the plane of polarization. *(b)* Exact cancellation of this rotation occurs if a molecule of *(S)*-2-butanol is encountered. *(c)* Net rotation of the plane of polarization occurs if (*R*)-2-butanol is present predominantly or exclusively.

In a racemic mixture the effect of each molecule of one enantiomer on the circularly-polarized beam cancels the effect of molecules of the other enantiomer, resulting in no net optical activity.

The racemic form of a sample is often designated as being (±). A racemic mixture of (*R*)-(−)-2-butanol and (*S*)-(+)-2-butanol might be indicated as

$$(\pm)\text{-2-butanol} \quad \text{or} \quad (\pm)\text{-}CH_3CH_2CHOHCH_3$$

5.9B Racemic Forms and Enantiomeric Excess

A sample of an optically active substance that consists of a single enantiomer is said to be **enantiomerically pure** or to have an **enantiomeric excess** of 100%. An enantiomerically pure sample of (*S*)-(+)-2-butanol shows a specific rotation of +13.52 ($[\alpha]_D^{25} = +13.52$). On the other hand, a sample of (*S*)-(+)-2-butanol that contains less than an equimolar amount of (*R*)-(−)-2-butanol will show a specific rotation that is less than +13.52 but greater than zero. Such a sample is said to have an *enantiomeric excess* less than 100%. The **enantiomeric excess (ee)**, also known as the **optical purity**, is defined as follows:

$$\%\ \text{Enantiomeric excess} = \frac{\text{moles of one enantiomer} - \text{moles of other enantiomer}}{\text{total moles of both enantiomers}} \times 100$$

The enantiomeric excess can be calculated from optical rotations:

$$\%\ \text{Enantiomeric excess}^* = \frac{\text{observed specific rotation}}{\text{specific rotation of the pure enantiomer}} \times 100$$

Let us suppose, for example, that a mixture of the 2-butanol enantiomers showed a specific rotation of +6.76. We would then say that the enantiomeric excess of the (*S*)-(+)-2-butanol is 50%:

$$\text{Enantiomeric excess} = \frac{+6.76}{+13.52} \times 100 = 50\%$$

When we say that the enantiomeric excess of this mixture is 50%, we mean that 50% of the mixture consists of the (+) enantiomer (the excess) and the other 50% consists of the racemic form. Since for the 50% that is racemic the optical rotations cancel one another out, only the 50% of the mixture that consists of the (+) enantiomer contributes to the observed optical rotation. The observed rotation is, therefore, 50% (or one-half) of what it would have been if the mixture had consisted only of the (+) enantiomer.

Solved Problem 5.5

What is the actual stereoisomeric composition of the mixture referred to above?

ANSWER Of the total mixture, 50% consists of the racemic form, which contains equal numbers of the two enantiomers. Therefore, half of this 50%, or 25%, is the (−) enantiomer and 25% is the (+) enantiomer. The other 50% of the mixture (the excess) is also the (+) enantiomer. Consequently, the mixture is 75% (+) enantiomer and 25% (−) enantiomer.

Review Problem 5.15

A sample of 2-methyl-1-butanol (see Section 5.8C) has a specific rotation, $[\alpha]_D^{25}$, equal to +1.151. **(a)** What is the percent enantiomeric excess of the sample? **(b)** Which enantiomer is in excess, the (*R*) or the (*S*)?

5.10 *The Synthesis of Chiral Molecules*

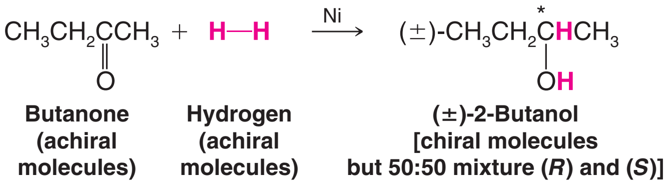
5.10A Racemic Forms

Reactions carried out with achiral reactants can often lead to *chiral* products. In the absence of any chiral influence from a catalyst, reagent, or solvent, the outcome of such a reaction is a racemic mixture. In other words, the chiral product is obtained as a 50:50 mixture of enantiomers.

---

*This calculation should be applied to a single enantiomer or to mixtures of enantiomers only. It should not be applied to mixtures in which some other compound is present.

An example is the synthesis of 2-butanol by the nickel-catalyzed hydrogenation of butanone. In this reaction the hydrogen molecule adds across the carbon–oxygen double bond in much the same way that it adds to a carbon–carbon double bond.



$$CH_3CH_2\underset{\underset{O}{\|}}{C}CH_3 + H{-}H \xrightarrow{Ni} (\pm)\text{-}CH_3CH_2\overset{*}{\underset{\underset{OH}{|}}{C}}HCH_3$$

Figure 5.15 illustrates the stereochemical aspects of this reaction. Because butanone is achiral, there is no difference in presentation of either face of the molecule to the surface of the metal catalyst. The two faces of the trigonal planar carbonyl group interact with the metal surface with equal probability. Transfer of the hydrogen atoms from the metal to the carbonyl group produces a chirality center at carbon 2. Since there has been no chiral influence in the reaction pathway, the product is obtained as a racemic mixture of the two enantiomers, (*R*)-(−)-2-butanol and (*S*)-(+)-2-butanol.

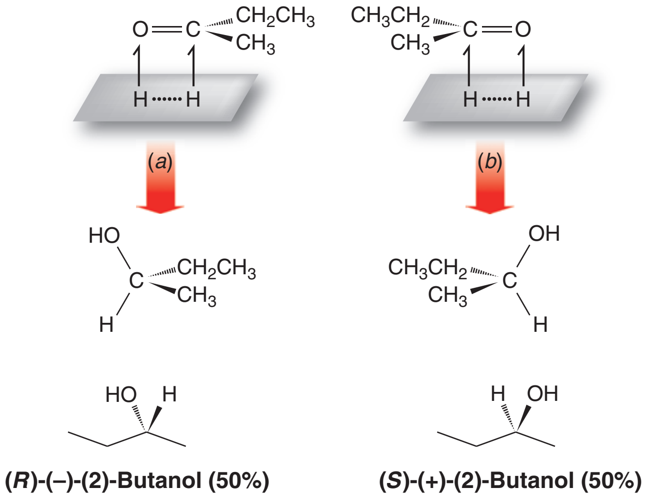


Figure 5.15 The reaction of butanone with hydrogen in the presence of a nickel catalyst. The reaction rate by path *(a)* is equal to that by path *(b)*. *(R-15)*-(−)-2-Butanol and *(S)*-(+)-2-butanol are produced in equal amounts, as a racemate.

We shall see that when reactions like this are carried out in the presence of a chiral influence, such as an enzyme or chiral catalyst, the result is usually not a racemic mixture.

5.10B Stereoselective Syntheses

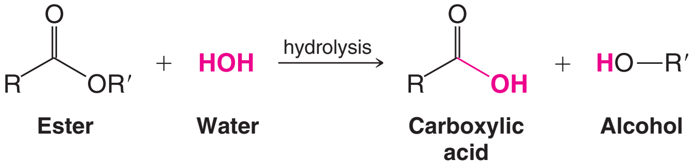
Stereoselective reactions are reactions that lead to a preferential formation of one stereoisomer over other stereoisomers that could possibly be formed.

- If a reaction produces preferentially one enantiomer over its mirror image, the reaction is said to be **enantioselective**.
- If a reaction leads preferentially to one diastereomer over others that are possible, the reaction is said to be **diastereoselective**.

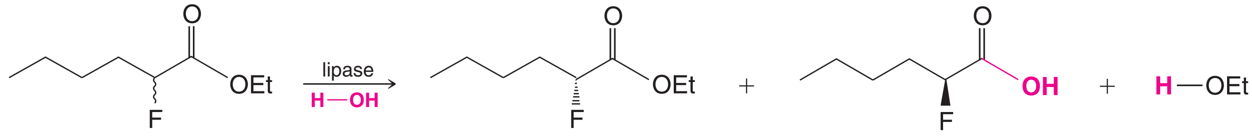
For a reaction to be either enantioselective or diastereoselective, a chiral reagent, catalyst, or solvent must assert an influence on the course of the reaction.

In nature, where most reactions are stereoselective, the chiral influences come from protein molecules called **enzymes**. Enzymes are biological catalysts of extraordinary efficiency. Not only do they have the ability to cause reactions to take place much more rapidly than they would otherwise, they also have the ability to assert a *dramatic chiral influence* on a reaction. Enzymes do this because they, too, are chiral, and they possess an active site where the reactant molecules are momentarily bound while the reaction takes place. The active site is chiral (See Fig. 5.7), and only one enantiomer of a chiral reactant fits it properly and is able to undergo the reaction.

Many enzymes have found use in the organic chemistry laboratory, where organic chemists take advantage of their properties to bring about stereoselective reactions. One of these is an enzyme called **lipase**. Lipase catalyzes a reaction called **hydrolysis**, whereby an ester (Section 2.10B) reacts with a molecule of water to produce a carboxylic acid and an alcohol.



If the starting ester is chiral and present as a mixture of its enantiomers, the lipase enzyme reacts selectively with one enantiomer to release the corresponding chiral carboxylic acid and an alcohol, while the other ester enantiomer remains unchanged or reacts much more slowly. The result is a mixture that consists predominantly of one stereoisomer of the reactant and one stereoisomer of the product, which can usually be separated easily on the basis of their different physical properties. Such a process is called a **kinetic resolution**, where the rate of a reaction with one enantiomer is different than with the other, leading to a preponderance of one product stereoisomer. We shall say more about the resolution of enantiomers in Section 5.16. The following hydrolysis is an example of a kinetic resolution using lipase:



Ethyl (±)-2-fluorohexanoate [an ester that is a racemate of (*R*) and (*S*) forms]

Ethyl (*R*)-(+)-2-fluorohexanoate (>99% enantiomeric excess)

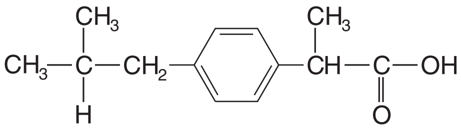
(*S*)-(−)-2-Fluorohexanoic acid (>69% enantiomeric excess)

Other enzymes called hydrogenases have been used to effect enantioselective versions of carbonyl reductions like that in Section 5.10A. We shall have more to say about the stereoselectivity of enzymes in Chapter 12.

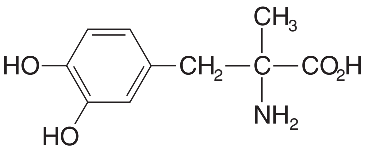
5.11 Chiral Drugs

The U.S. Food and Drug Administration and the pharmaceutical industry are very interested in the production of chiral drugs, that is, drugs that contain a single enantiomer rather than a racemate. The antihypertensive drug **methyldopa** (Aldomet), for example, owes its effect exclusively to the (*S*) isomer. In the case of **penicillamine**, the (*S*) isomer is a highly potent therapeutic agent for primary chronic arthritis, while the (*R*) isomer has no therapeutic action and is highly toxic. The anti-inflammatory agent **ibuprofen** (Advil, Motrin, Nuprin) is marketed as a racemate even though only the (*S*) enantiomer is the active agent. The (*R*) isomer of ibuprofen has no anti-inflammatory action and is slowly converted to the (*S*) isomer in the body. A formulation of ibuprofen based on solely the (*S*) isomer, however, would be more effective than the racemate.

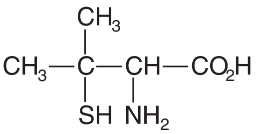
At the beginning of this chapter we showed the formulas for two enantiomeric drugs, Darvon and Novrad. Darvon (also called dextropropoxyphene) is a painkiller. Novrad (levopropoxyphene) is a cough suppressant.



Ibuprofen



Methyldopa



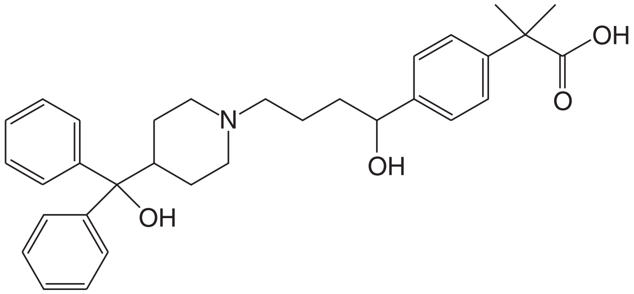
Penicillamine

Review Problem 5.16

Write three-dimensional formulas for the (*S*) isomers of **(a)** methyldopa, **(b)** penicillamine, and **(c)** ibuprofen.

Review Problem 5.17

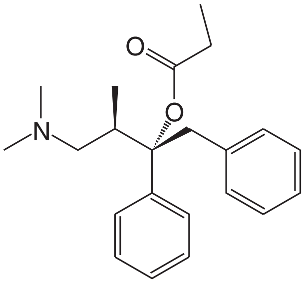
The antihistamine Allegra (fexofenadine) has the following structural formula. For any chirality centers in fexofenadine, draw a substructure that would have an (*R*) configuration.



Fexofenadine (Allegra)

Review Problem 5.18

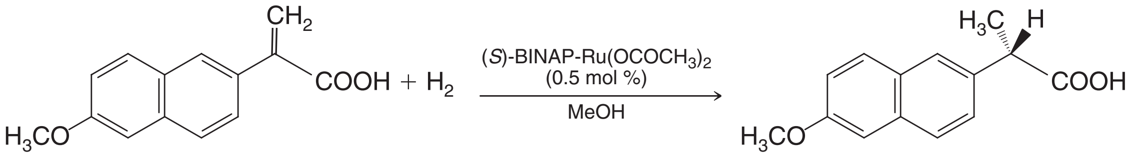
Assign the (*R,S*) configuration at each chirality center in Darvon (dextropropoxyphene).



Darvon

There are many other examples of drugs like these, including drugs where the enantiomers have distinctly different effects. The preparation of enantiomerically pure drugs, therefore, is one factor that makes stereoselective synthesis (Section 5.10B) and the resolution of racemic drugs (separation into pure enantiomers, Section 5.16) major areas of research today.

Underscoring the importance of stereoselective synthesis is the fact that the 2001 Nobel Prize in Chemistry was given to researchers who developed reaction catalysts that are now widely used in industry and academia. William Knowles (Monsanto Company, retired) and Ryoji Noyori (Nagoya University) were awarded half of the prize for their development of reagents used for catalytic stereoselective hydrogenation reactions. The other half of the prize was awarded to Barry Sharpless (Scripps Research Institute) for development of catalytic stereoselective oxidation reactions (see Chapter 11). An important example resulting from the work of Noyori and based on earlier work by Knowles is a synthesis of the anti-inflammatory agent **naproxen,** involving a stereoselective catalytic hydrogenation reaction:



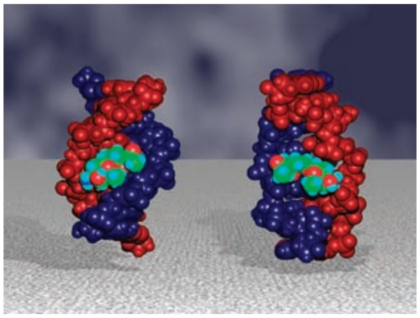
(*S*)-Naproxen
(an anti-inflammatory agent)
(92% yield, 97% ee)

The hydrogenation catalyst in this reaction is an organometallic complex formed from ruthenium and a chiral organic ligand called (*S*)-BINAP. The reaction itself is truly remarkable because it proceeds with excellent enantiomeric excess (97%) and in very high yield (92%). We will have more to say about BINAP ligands and the origin of their chirality in Section 5.18.

THE CHEMISTRY OF . . .

Selective Binding of Drug Enantiomers to Left- and Right-Handed Coiled DNA

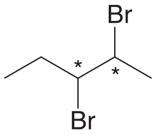
Would you like left- or right-handed DNA with your drug? That's a question that can now be answered due to the recent discovery that each enantiomer of the drug daunorubicin selectively binds DNA coiled with opposite handedness. (+)-Daunorubicin binds selectively to DNA coiled in the typical right-handed conformation (B-DNA). (−)-Daunorubicin binds selectively to DNA coiled in the left-handed conformation (Z-DNA). Furthermore, daunorubicin is capable of inducing conformational changes in DNA from one coiling direction to the other, depending on which coiling form is favored when a given daunorubicin enantiomer binds to the DNA. It has long been known that DNA adopts a number of secondary and tertiary structures, and it is presumed that some of these conformations are involved in turning on or off transcription of a given section of DNA. The discovery of specific interactions between each daunorubicin enantiomer and the left- and right-handed coil forms of DNA will likely assist in design and discovery of new drugs with anticancer or other activities.



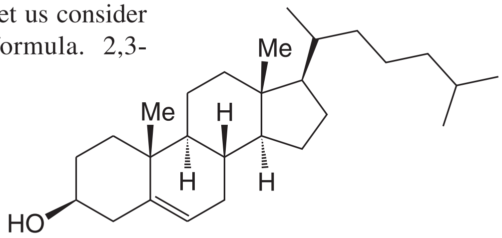
Enantiomeric forms of daunorubicin bind with DNA and cause it to coil with opposite handedness. [Graphic courtesy John O. Trent, Brown Cancer Center, Department of Medicine, University of Louisville, KY. Based on work from Qu, X. Trent, J.O., Fokt, I., Priebe, W., and Chaires, J.B., *Allosteric, Chiral-Selective Drug Building to DNA, Proc. Natl. Acad. Sci. U.S.A.*, 2000 (Oct. 24): 97(22), 12032–12037.]

5.12 *Molecules with More than One Chirality Center*

So far we have mainly considered chiral molecules that contain only one chirality center. Many organic molecules, especially those important in biology, contain more than one chirality center. Cholesterol (Section 23.4B), for example, contains eight chirality centers. (Can you locate them?) We can begin, however, with simpler molecules. Let us consider 2,3-dibromopentane, shown here in a two-dimensional bond-line formula. 2,3-Dibromopentane has two chirality centers:



2,3-Dibromopentane



Cholesterol

A useful rule gives the maximum number of stereoisomers:

- In compounds whose stereoisomerism is due to chirality centers, ***the total number of stereoisomers will not exceed 2^n, where n is equal to the number of chirality centers***.

Helpful Hint

Cholesterol, having eight chirality centers, hypothetically could exist in 2^8 (256) stereoisomeric forms, yet biosynthesis via enzymes produces only *one* stereoisomer.

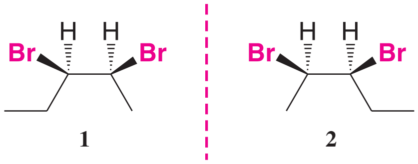
For 2,3-dibromopentane we should not expect more than four stereoisomers ($2^2 = 4$).

Our next task is to write three-dimensional bond-line formulas for the possible stereoisomers. When doing so it is helpful to follow certain conventions. First, it is generally best to write as many carbon atoms in the plane of the paper as possible. Second, when needing to compare the stereochemistry at adjacent carbon atoms, we usually draw the molecule in a fashion that shows eclipsing interactions, even though this would not be the most stable conformation of the molecule. We do so because, as we shall see later, eclipsed conformations make it easy for us to recognize planes of symmetry when they are present. (We do not mean to imply, however, that eclipsed conformations are the most stable ones—they most certainly are not. It is important to remember that free rotation is possible about single bonds, and that molecules are constantly changing conformations.) Third, if we need to draw the enantiomer

Helpful Hint

Useful conventions when writing three-dimensional formulas of the first stereoisomer, we can easily do so by drawing a mirror image of the first formula, using as our guide an imaginary mirror perpendicular to the page and *between* the molecules.

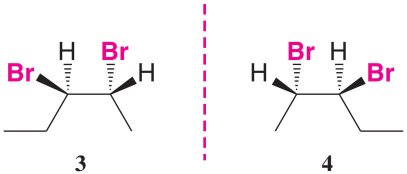
The following are two three-dimensional bond-line formulas for 2,3-dibromopentane. Notice that in drawing these formulas we have followed the conventions above.



1 **2**

Since structures **1** and **2** are not superposable, they represent different compounds. Since structures **1** and **2** differ only in the arrangement of their atoms in space, they represent stereoisomers. Structures **1** and **2** are also mirror images of each other; thus **1** and **2** represent a pair of enantiomers.

Structures **1** and **2** are not the only ones possible for 2,3-dibromopentane, however. If we interchange the bromine and hydrogen at C2 (invert the configuration), we find that we have **3**, which has a different structural formula than either **1** or **2**. Furthermore, we can write a formula for a structure (**4**) that is a nonsuperposable mirror image of **3**, and which is also different from **1** and **2**.



3 **4**

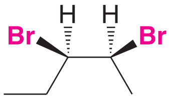
Structures **3** and **4** correspond to another pair of enantiomers. Structures **1–4** are all different, so there are, in total, four stereoisomers of 2,3-dibromopentane. Essentially what we have done above is to write all the possible structures that result by successively interchanging two groups at all chirality centers. At this point you should convince yourself that there are no other stereoisomers by writing other structural formulas. You will find that rotation about the single bonds (or of the entire structure) of any other arrangement of the atoms will cause the structure to become superposable with one of the structures that we have written here. Better yet, using different colored balls, make molecular models as you work this out.

The compounds represented by structures **1–4** are all optically active compounds. Any one of them, if placed separately in a polarimeter, would show optical activity.

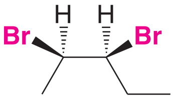
The compounds represented by structures **1** and **2** are enantiomers. The compounds represented by structures **3** and **4** are also enantiomers. But what is the isomeric relation between the compounds represented by **1** and **3**?

We can answer this question by observing that **1** and **3** *are stereoisomers* and that they *are not mirror images of each other*. They are, therefore, *diastereomers*.

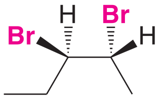
- Diastereomers have different physical properties—different melting points and boiling points, different solubilities, and so forth.



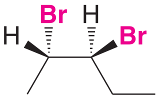
1



2



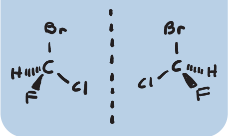
3



4

Review Problem 5.19

(a) If **3** and **4** are enantiomers, what are **1** and **4**? **(b)** What are **2** and **3**, and **2** and **4**? **(c)** Would you expect **1** and **3** to have the same melting point? **(d)** The same boiling point? **(e)** The same vapor pressure?

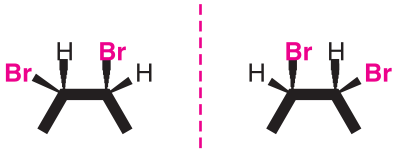


5.12A Meso Compounds

A structure with two chirality centers does not always have four possible stereoisomers. Sometimes there are only *three*. As we shall see:

- Some molecules are achiral even though they contain chirality centers.

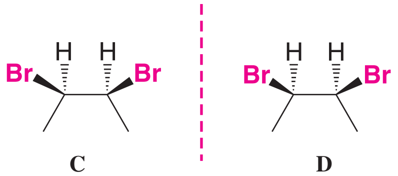
To understand this, let us write stereochemical formulas for 2,3-dibromobutane. We begin in the same way as we did before. We write formulas for one stereoisomer and for its mirror image:



A **B**

Structures **A** and **B** are nonsuperposable and represent a pair of enantiomers.

When we write the new structure **C** (see below) and its mirror image **D**, however, the situation is different. *The two structures are superposable*. This means that **C** and **D** do not represent a pair of enantiomers. Formulas **C** and **D** represent identical orientations of the same compound:



C **D**

The molecule represented by structure **C** (or **D**) is not chiral even though it contains two chirality centers.

- A **meso compound** is an achiral molecule that contains chirality centers. Meso compounds are not optically active.

The ultimate test for molecular chirality is to construct a model (or write the structure) of the molecule and then test whether or not the model (or structure) is superposable on its mirror image. If it is, the molecule is achiral: If it *is not*, the molecule is chiral.

We have already carried out this test with structure **C** and found that it is achiral. We can also demonstrate that **C** is achiral in another way. Figure 5.16 shows that structure **C** *has an internal plane of symmetry* (Section 5.6).

The following two problems relate to compounds **A–D** in the preceding paragraphs.

Helpful Hint

If a molecule has an internal plane of symmetry it is achiral.

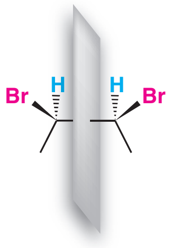


Figure 5.16 The plane of symmetry of *meso*-2,3-dibromobutane. This plane divides the molecule into halves that are mirror images of each other.

Review Problem 5.20

Which of the following would be optically active?

(a) A pure sample of **A**

(b) A pure sample of **B**

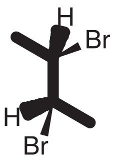
(c) A pure sample of **C**

(d) An equimolar mixture of **A** and **B**

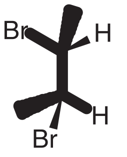
Review Problem 5.21

The following are formulas for three compounds, written in noneclipsed conformations. In each instance tell which compound (**A**, **B**, or **C** above) each formula represents.

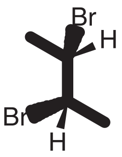
(a)



(b)



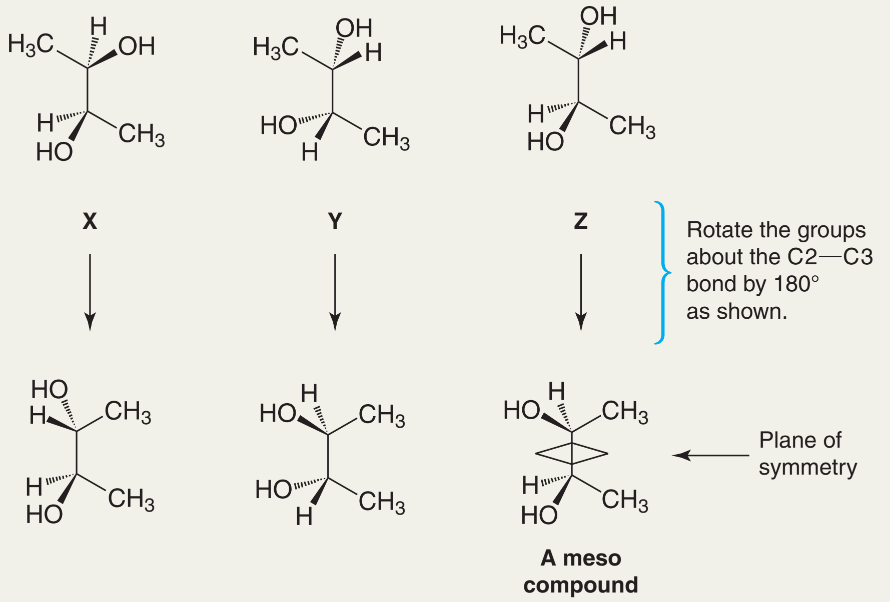
(c)



Solved Problem 5.6

Which of the following is a meso compound?

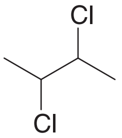
STRATEGY AND ANSWER In each molecule, rotating the groups joined by the C2—C3 bond by 180° brings the two methyl groups into comparable position. In the case of compound Z, a plane of symmetry results, and therefore, Z is a meso compound. No plane of symmetry is possible in X and Y.



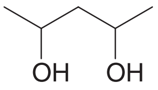
Review Problem 5.22

Write three-dimensional formulas for all of the stereoisomers of each of the following compounds. Label pairs of enantiomers and label meso compounds.

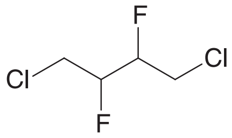
(a)



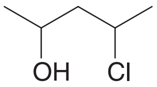
(b)



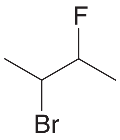
(c)



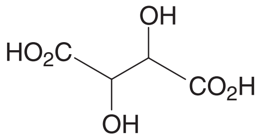
(d)



(e)



(f)

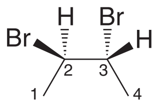


Tartaric acid

5.12B How to Name Compounds with More than One Chirality Center

If a compound has more than one chirality center, we analyze each center separately and decide whether it is (*R*) or (*S*). Then, using numbers, we tell which designation refers to which carbon atom.

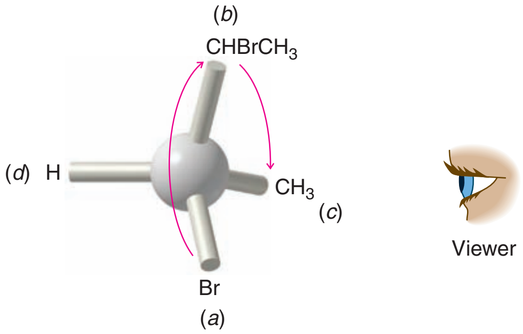
Consider stereoisomer **A** of 2,3-dibromobutane:



A

2,3-Dibromobutane

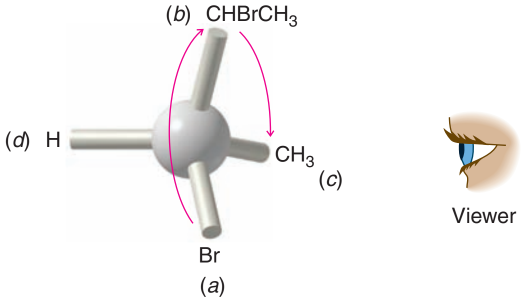
When this formula is rotated so that the group of lowest priority attached to C2 is directed away from the viewer, it resembles the following:



(*R*) Configuration

The order of progression from the group of highest priority to that of next highest priority (from —Br, to $—CHBrCH_3$, to $—CH_3$) is clockwise. So C2 has the (*R*) configuration.

When we repeat this procedure with C3, we find that C3 also has the (*R*) configuration:



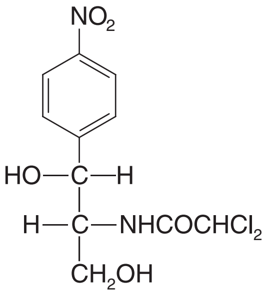
(*R*) Configuration

Compound **A**, therefore, is (2*R*,3*R*)-2,3-dibromobutane.

Give names that include (*R*) and (*S*) designations for compounds **B** and **C** in Section 5.12A. *Review Problem 5.23*

Give names that include (*R*) and (*S*) designations for your answers to Review Problem 5.22. *Review Problem 5.24*

Chloramphenicol (at right) is a potent antibiotic, isolated from *Streptomyces venezuelae*, that is particularly effective against typhoid fever. It was the first naturally occurring substance shown to contain a nitro ($—NO_2$) group attached to an aromatic ring. Both chirality centers in chloramphenicol are known to have the (*R*) configuration. Identify the two chirality centers and write a three-dimensional formula for chloramphenicol. *Review Problem 5.25*



Chloramphenicol

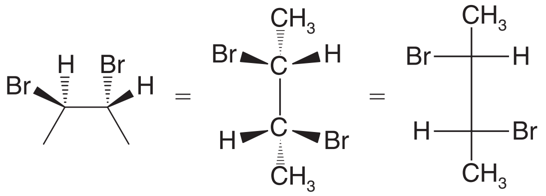
5.13 *Fischer Projection Formulas*

So far in writing structures for chiral molecules we have only used formulas that show three dimensions with solid and dashed wedges, and we shall largely continue to do so until we study carbohydrates in Chapter 22. The reason is that formulas with solid and dashed wedges unambiguously show three dimensions, and they can be manipulated on paper in any way that we wish so long as we do not break bonds. Their use, moreover, teaches us to see molecules (in our mind's eye) in three dimensions, and this ability will serve us well.

Chemists, however, sometimes use formulas called **Fischer projections** to show three dimensions in chiral molecules such as acyclic carbohydrates. Fischer projection formulas are useful in cases where there are chirality centers at several adjacent carbon atoms, as is often the case in carbohydrates. Use of Fischer projection formulas requires rigid adherence to certain conventions, however. **Used carelessly, these projection formulas can easily lead to incorrect conclusions**.

5.13A How to Draw and Use Fischer Projections

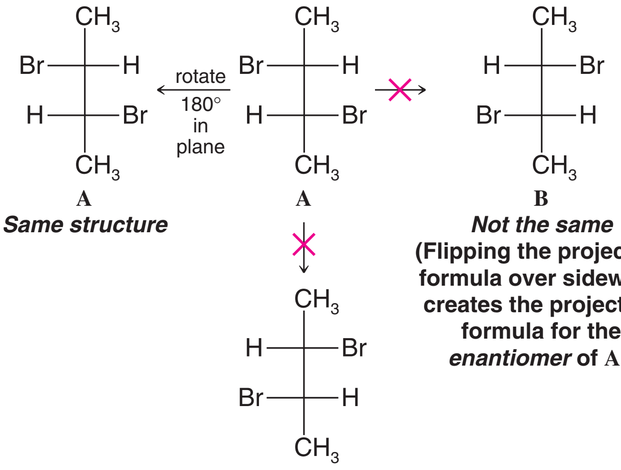
Let us consider how we would relate a three-dimensional formula for 2,3-dibromobutane using solid and dashed wedges to the corresponding Fischer projection formula. First, it is necessary to note that in Fischer projections the carbon chain is always drawn from top to bottom, rather than side to side as is often the case with bond-line formulas. We consider the molecule in a conformation that has eclipsing interactions between the groups at each carbon. For 2,3-dibromobutane we turn the bond-line formula so that the carbon chain runs up and down and we orient it so that groups attached to the main carbon chain project out of the plane like a bow tie. The carbon–carbon bonds of the chain, therefore, lie either in the plane of the paper or project behind it. Then to draw the Fischer projection we simply "project" all of the bonds onto the paper, replacing all solid and dashed wedges with ordinary lines. Having done this, the vertical line of the formula now represents the carbon chain, each point of intersection between the vertical line and a horizontal line represents a carbon atom in the chain, and we understand the horizontal lines to be bonds that project out toward us.



A **A** **A**

Fischer projection formula

To test the superposability of two structures represented by Fischer projections we are allowed to rotate them in the plane of the paper by 180°, *but by no other angle*. We must always keep the Fischer projection formulas in the plane of the paper, and **we are not allowed to flip them over**. If we flip a Fischer projection over, the horizontal bonds project behind the plane instead of in front, and every configuration would be *misrepresented* as the opposite of what was intended.



A
Same structure

A

B
Not the same
(Flipping the projection formula over sideways creates the projection formula for the *enantiomer* of A.)

Not the same
(Flipping the projection formula over end for end creates the projection formula for the *enantiomer* of A.)

Helpful Hint

Build handheld models of **A** and **B** and relate them to the Fischer projections shown here.

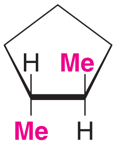
Because Fischer projections must be used with such care, we introduce them now only so that you can understand Fischer projections when you see them in the context of other courses. Our emphasis for most of this book will be on the use of solid and dashed wedges to represent three-dimensional formulas (or chair conformational structures in the case of cyclohexane derivatives), except in Chapter 22 when we will use Fischer projections again in our discussion of carbohydrates. If your instructor wishes to utilize Fischer projections further, you will be so advised.

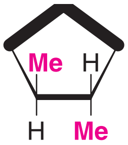
Review Problem 5.26

(**a**) Give the (*R,S*) designations for each chirality center in compound **A** and for compound **B**. (**b**) Write the Fischer projection formula for a compound **C** that is the diastereomer of **A** and **B**. (**c**) Would **C** be optically active?

5.14 *Stereoisomerism of Cyclic Compounds*

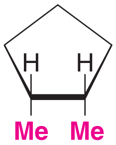
Cyclopentane derivatives offer a convenient starting point for a discussion of the stereoisomerism of cyclic compounds. For example, 1,2-dimethylcyclopentane has two chirality centers and exists in three stereoisomeric forms **5**, **6**, and **7**:





Enantiomers

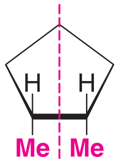
5 **6**



Meso compound

7

The trans compound exists as a pair of enantiomers **5** and **6**. *cis*-1,2-Dimethylcyclopentane (**7**) is a meso compound. It has a plane of symmetry that is perpendicular to the plane of the ring:



Plane of symmetry

7

Review Problem 5.27

(a) Is *trans*-1,2-dimethylcyclopentane (**5**) superposable on its mirror image (i.e., on compound **6**)? **(b)** Is *cis*-1,2-dimethylcyclopentane (**7**) superposable on its mirror image? **(c)** Is *cis*-1,2-dimethylcyclopentane a chiral molecule? **(d)** Would *cis*-1,2-dimethylcyclopentane show optical activity? **(e)** What is the stereoisomeric relationship between **5** and **7**? **(f)** Between **6** and **7**?

Review Problem 5.28

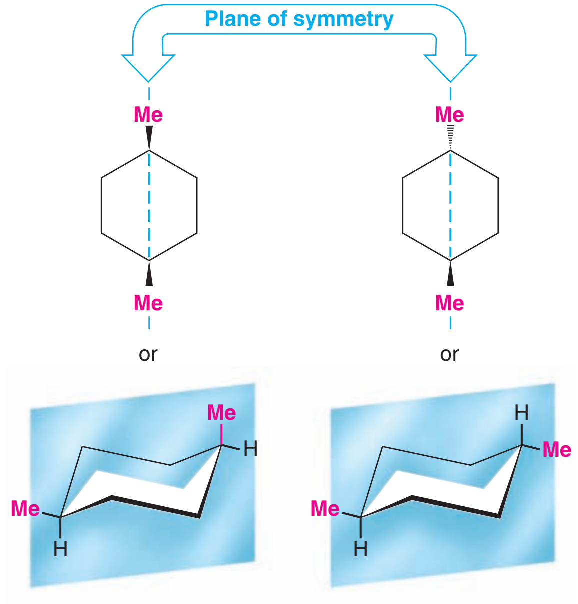
Write structural formulas for all of the stereoisomers of 1,3-dimethylcyclopentane. Label pairs of enantiomers and meso compounds if they exist.

5.14A Cyclohexane Derivatives

1,4-Dimethylcyclohexanes If we examine a formula of 1,4-dimethylcyclohexane, we find that it does not contain any chirality centers. However, it does have two stereogenic centers. As we learned in Section 4.13, 1,4-dimethylcyclohexane can exist as cis–trans isomers. The cis and trans forms (Fig. 5.17) are *diastereomers*. Neither compound is chiral and, therefore, neither is optically active. Notice that both the cis and trans forms of 1,4-dimethylcyclohexane have a plane of symmetry.

Helpful Hint

Build handheld molecular models of the 1,4-, 1,3-, and 1,2-dimethylcyclohexane isomers discussed here and examine their stereochemical properties. Experiment with flipping the chairs and also switching between cis and trans isomers.



***cis*-1,4-Dimethylcyclohexane** ***trans*-1,4-Dimethylcyclohexane**

Figure 5.17 The cis and trans forms of 1,4-dimethylcyclohexane are diastereomers of each other. Both compounds are achiral, as the internal plane of symmetry (blue) shows for each.

1,3-Dimethylcyclohexanes 1,3-Dimethylcyclohexane has two chirality centers; we can, therefore, expect as many as four stereoisomers ($2^2 = 4$). In reality there are only three. *cis*-1,3-Dimethylcyclohexane has a plane of symmetry (Fig. 5.18) and is achiral.

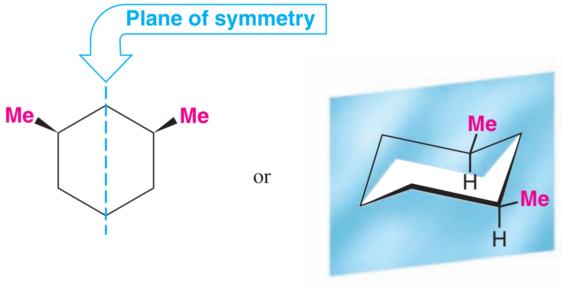
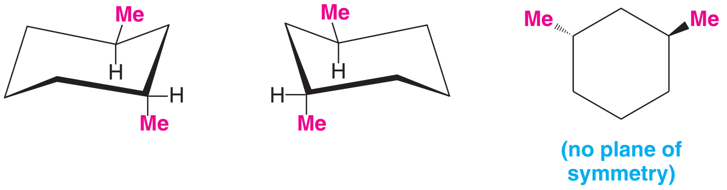


Figure 5.18 *cis*-1,3-Dimethylcyclohexane has a plane of symmetry, shown in blue, and is therefore achiral.

trans-1,3-Dimethylcyclohexane does not have a plane of symmetry and exists as a pair of enantiomers (Fig. 5.19). You may want to make models of the *trans*-1,3-dimethylcyclohexane enantiomers. Having done so, convince yourself that they cannot be superposed as they stand and that they cannot be superposed after one enantiomer has undergone a ring flip.



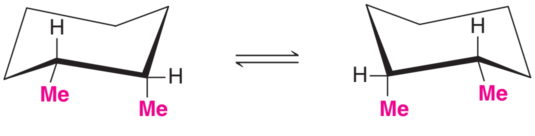
(*a*) (*b*) (*c*)

Figure 5.19 *trans*-1,3-Dimethylcyclohexane does not have a plane of symmetry and exists as a pair of enantiomers. The two structures *(a* and *b)* shown here are not superposable as they stand, and flipping the ring of either structure does not make it superposable on the other. *(c)* A simplified representation of *(b)*.

1,2-Dimethylcyclohexanes 1,2-Dimethylcyclohexane also has two chirality centers, and again we might expect as many as four stereoisomers. Indeed there are four, but we find that we can *isolate* only *three* stereoisomers. *trans*-1,2-Dimethylcyclohexane (Fig. 5.20) exists as a pair of enantiomers. Its molecules do not have a plane of symmetry.

(*a*) (*b*)

Figure 5.20 *trans*-1,2-Dimethylcyclohexane has no plane of symmetry and exists as a pair of enantiomers (*a* and *b*). [Notice that we have written the most stable conformations for *(a)* and *(b)*. A ring flip of either *(a)* or *(b)* would cause both methyl groups to become axial.]



(*c*) (*d*)

Figure 5.21 *cis*-1,2-Dimethylcyclohexane exists as two rapidly interconverting chair conformations (*c*) and (*d*).

cis-1,2-Dimethylcyclohexane, shown in Fig. 5.21, presents a somewhat more complex situation. If we consider the two conformational structures (*c*) and (*d*), we find that these two mirror-image structures are not identical. Neither has a plane of symmetry and each is a chiral molecule, *but they are interconvertible by a ring flip*. Therefore, although the two structures represent enantiomers, *they cannot be separated* because they rapidly interconvert even at low temperature. They simply represent *different conformations of the same compound*. Therefore, structures (*c*) and (*d*) are not configurational stereoisomers; they are **conformational stereoisomers** (see Section 4.9A). This means that at normal temperatures there are only three *isolable stereoisomers* of 1,2-dimethylcyclohexane.

As we shall see later, there are some compounds whose conformational stereoisomers *can* be isolated in enantiomeric forms. Isomers of this type are called atropisomers (Section 5.18).

Review Problem 5.29

Write formulas for all of the isomers of each of the following. Designate pairs of enantiomers and achiral compounds where they exist.

(a) 1-Bromo-2-chlorocyclohexane **(b)** 1-Bromo-3-chlorocyclohexane

(c) 1-Bromo-4-chlorocyclohexane

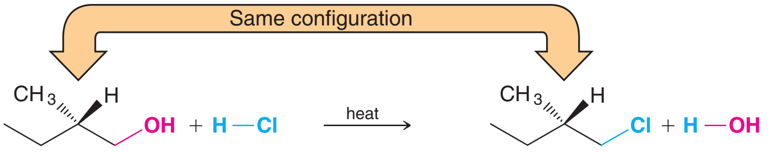
Review Problem 5.30

Give the (*R*,*S*) designation for each compound given as an answer to Review Problem 5.29.

5.15 *Relating Configurations through Reactions in Which No Bonds to the Chirality Center Are Broken*

- **If a reaction takes place in a way so that no bonds to the chirality center are broken**, the product will of necessity have the same general configuration of groups around the chirality center as the reactant.

Such a reaction is said **to proceed with retention of configuration**. Consider as an example the reaction that takes place when (*S*)-(−)-2-methyl-1-butanol is heated with concentrated hydrochloric acid:

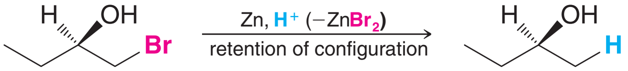


(*S*)-(−)-2-Methyl-1-butanol **$[\alpha]_D^{25} = -5.756$**

(*S*)-(+)-1-Chloro-2-methylbutane **$[\alpha]_D^{25} = +1.64$**

We do not need to know now exactly how this reaction takes place to see that the reaction must involve breaking the CH_2—OH bond of the alcohol because the —OH group is replaced by a —Cl. There is no reason to assume that any other bonds are broken. (We shall study how this reaction takes place in Section 11.8A.) Since no bonds to the chirality center are broken, the reaction must take place with retention of configuration, and the product of the reaction *must have the same configuration of groups around the chirality center that the reactant had*. By saying that the two compounds have the same configuration, we simply mean that comparable or identical groups in the two compounds occupy the same relative positions in space around the chirality center. (In this instance the —CH_2OH group and the —CH_2Cl are comparable, and they occupy the same relative position in both compounds; all the other groups are identical and they occupy the same positions.)

Notice that in this example while the (*R,S*) designation *does not change* [both reactant and product are (*S*)], the direction of optical rotation *does change* [the reactant is (−) and the product is (+)]. Neither occurrence is a necessity when a reaction proceeds with retention of configuration. In the next section we shall see examples of reactions in which configurations are retained and where the direction of optical rotation does not change. The following reaction is one that proceeds with retention of configuration but involves a change in the (*R,S*) designation:



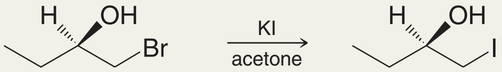
(*R*)-1-Bromo-2-butanol **(*S*)-2-Butanol**

In this example the (*R,S*) designation changes because the —CH_2Br group of the reactant changes to a —CH_3 group in the product (—CH_2Br has a higher priority than —CH_2CH_3, and —CH_3 has a lower priority than —CH_2CH_3).

Solved Problem 5.7

When (*R*)-1-bromo-2-butanol reacts with KI in acetone the product is 1-iodo-2-butanol. Would the product be (*R*) or (*S*)?

STRATEGY AND ANSWER No bonds to the chirality center would be broken, so we can reason that the product would be the following.



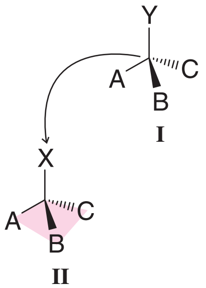
(R)-1-Bromo-2-butanol **(R)-1-Iodo-2-butanol**

The configuration of the product would still be (*R*) because replacing the bromine at C1 with an iodine atom does not change the relative priority of C1.

5.15A Relative and Absolute Configurations

Reactions in which no bonds to the chirality center are broken are useful in relating configurations of chiral molecules. That is, they allow us to demonstrate that certain compounds have the same relative configuration. In each of the examples that we have just cited, the products of the reactions have the same *relative configurations* as the reactants.

- Chirality centers in different molecules have the same **relative configuration** if they share three groups in common and if these groups **with** the central carbon can be superposed in a pyramidal arrangement.

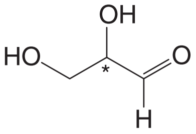


The chirality centers in **I** and **II** have the same relative configuration. Their common groups and central carbon can be superposed.

Before 1951 only relative configurations of chiral molecules were known. No one prior to that time had been able to demonstrate with certainty what the actual spatial arrangement of groups was in any chiral molecule. To say this another way, no one had been able to determine the absolute configuration of an optically active compound.

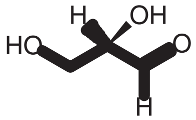
- The **absolute configuration** of a chirality center is its (*R*) or (*S*) designation, which can only be specified by knowledge of the actual arrangement of groups in space at the chirality center.

Prior to any known absolute configurations, the configurations of chiral molecules were related to each other *through reactions of known stereochemistry*. Attempts were also made to relate all configurations to a single compound that had been chosen arbitrarily to be the standard. This standard compound was glyceraldehyde:

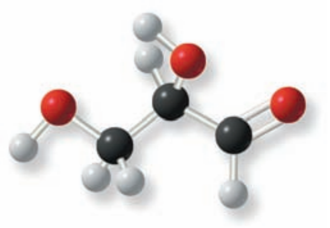


Glyceraldehyde

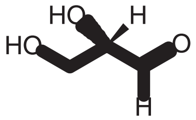
Glyceraldehyde has one chirality center; therefore, glyceraldehyde exists as a pair of enantiomers:



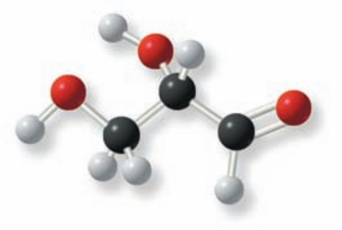
(*R*)-Glyceraldehyde



and



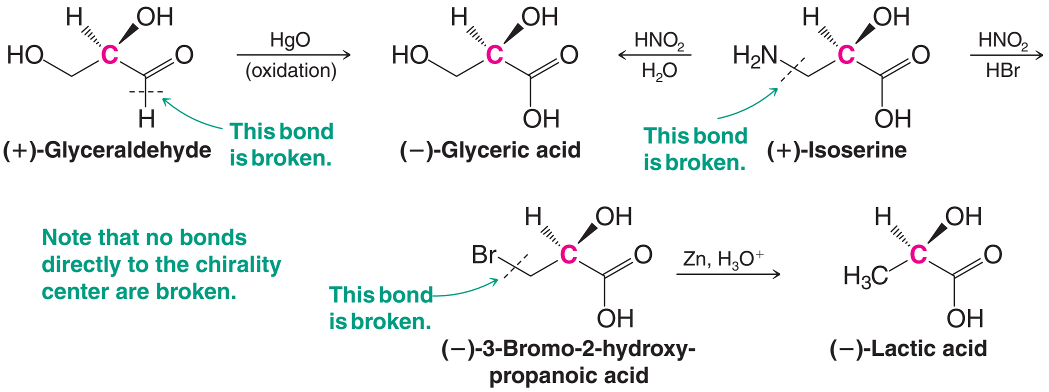
(*S*)-Glyceraldehyde



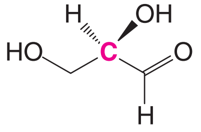
In one system for designating configurations, (*R*)-glyceraldehyde is called D-glyceraldehyde and (*S*)-glyceraldehyde is called L-glyceraldehyde. This system of nomenclature is used with a specialized meaning in the nomenclature of carbohydrates. (See Section 22.2B.)

One glyceraldehyde enantiomer is dextrorotatory (+) and the other, of course, is levorotatory (−). Before 1951 no one was sure, however, which configuration belonged to which enantiomer. Chemists decided arbitrarily to assign the (*R*) configuration to the (+)-enantiomer. Then, configurations of other molecules were related to one glyceraldehyde enantiomer or the other through reactions of known stereochemistry.

For example, the configuration of (−)-lactic acid can be related to (+)-glyceraldehyde through the following sequence of reactions in which no bond to the chirality center is broken:

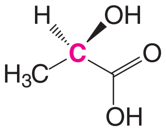


The stereochemistry of all of these reactions is known. Because none of the bonds to the chirality center (shown in red) has been broken during the sequence, its original configuration is retained. If the assumption is made that (+)-glyceraldehyde is the (*R*) stereoisomer, and therefore has the following configuration,



(*R*)-(+)-Glyceraldehyde

then (−)-lactic acid is also an (*R*) stereoisomer and its configuration is

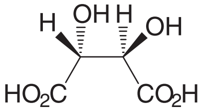


(*R*)-(−)-Lactic acid

Review Problem 5.31

Write bond-line three-dimensional formulas for the starting compound, the product, and all of the intermediates in a synthesis similar to the one just given that relates the configuration of (−)-glyceraldehyde with (+)-lactic acid. Label each compound with its proper (*R*) or (*S*) and (+) or (−) designation.

The configuration of (−)-glyceraldehyde was also related through reactions of known stereochemistry to (+)-tartaric acid:



(+)-Tartaric acid

In 1951 J. M. Bijvoet, the director of the van't Hoff Laboratory of the University of Utrecht in the Netherlands, using a special technique of X-ray diffraction, was able to show conclusively that (+)-tartaric acid had the absolute configuration shown above. This meant that the original arbitrary assignment of the configurations of (+)- and (−)-glyceraldehyde was also correct. It also meant that the configurations of all of the compounds that had been related to one glyceraldehyde enantiomer or the other were now known with certainty and were now **absolute configurations**.

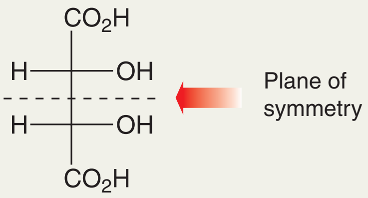
Review Problem 5.32

Fischer projection formulas are often used to depict compounds such as glyceraldehyde, lactic acid, and tartaric acid. Draw Fischer projections for both enantiomers of (**a**) glyceraldehyde, (**b**) tartaric acid, and (**c**) lactic acid, and specify the (*R*) or (*S*) configuration at each chirality center. [Note that in Fischer projection formulas the terminal carbon that is most highly oxidized is placed at the top of the formula (an aldehyde or carboxylic acid group in the specific examples here).]

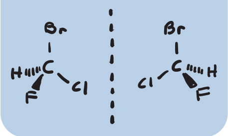
Solved Problem 5.8

Write a Fischer projection formula for a tartaric acid isomer that is not chiral.

STRATEGY AND ANSWER We reason that because tartaric acid has two chirality centers, the achiral isomer must have a plane of symmetry and be a meso compound.



***meso*-Tartaric acid**
(achiral)



5.16 *Separation of Enantiomers: Resolution*

So far we have left unanswered an important question about optically active compounds and racemic forms: How are enantiomers separated? Enantiomers have identical solubilities in ordinary solvents, and they have identical boiling points. Consequently, the conventional methods for separating organic compounds, such as crystallization and distillation, fail when applied to a racemic form.

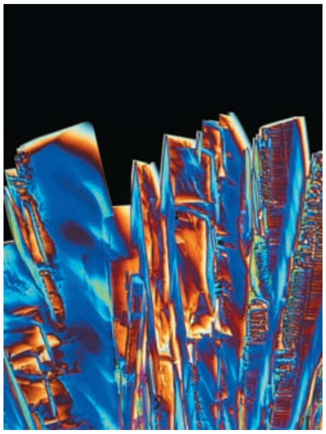
5.16A Pasteur's Method for Separating Enantiomers

It was, in fact, Louis Pasteur's separation of a racemic form of a salt of tartaric acid in 1848 that led to the discovery of the phenomenon called enantiomerism. Pasteur, consequently, is often considered to be the founder of the field of stereochemistry.

(+)-Tartaric acid is one of the by-products of wine making (nature usually only synthesizes one enantiomer of a chiral molecule). Pasteur had obtained a sample of racemic tartaric acid from the owner of a chemical plant. In the course of his investigation Pasteur began examining the crystal structure of the sodium ammonium salt of racemic tartaric acid. He noticed that two types of crystals were present. One was identical with crystals of the sodium ammonium salt of (+)-tartaric acid that had been discovered earlier and had been shown to be dextrorotatory. Crystals of the other type were *non*superposable mirror images of the first kind. The two types of crystals were actually chiral. Using tweezers and a magnifying glass, Pasteur separated the two kinds of crystals, dissolved them in water, and placed the solutions in a polarimeter. The solution of crystals of the first type was dextrorotatory, and the crystals themselves proved to be identical with the sodium ammonium salt of (+)-tartaric acid that was already known. The solution of crystals of the second type was levorotatory; it rotated plane-polarized light in the opposite direction and by an equal amount. The crystals of the second type were of the sodium ammonium salt of (−)-tartaric acid. The chirality of the crystals themselves disappeared, of course, as the crystals dissolved into their solutions, *but the optical activity* remained. Pasteur reasoned, therefore, that the molecules themselves must be chiral.

Pasteur's discovery of enantiomerism and his demonstration that the optical activity of the two forms of tartaric acid was a property of the molecules themselves led, in 1874, to the proposal of the tetrahedral structure of carbon by van't Hoff and Le Bel.

Unfortunately, few organic compounds give chiral crystals as do the (+)- and (−)-tartaric acid salts. Few organic compounds crystallize into separate crystals (containing separate enantiomers) that are visibly chiral like the crystals of the sodium ammonium salt of tartaric acid. Pasteur's method, therefore, is not generally applicable to the separation of enantiomers.



Tartaric acid crystals

5.16B Current Methods for Resolution of Enantiomers

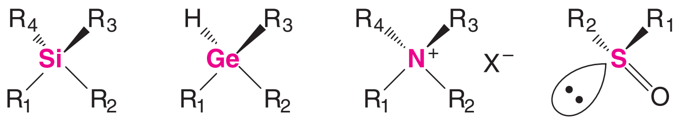
One of the most useful procedures for separating enantiomers is based on the following:

- When a racemic mixture reacts with a single enantiomer of another compound, a mixture of diastereomers results, and diastereomers, because they have different melting points, boiling points, and solubilities, can be separated by conventional means.

Diastereomeric recrystallization is one such process. We shall see how this is done in Section 20.3F. Another method is **resolution** by enzymes, whereby an enzyme selectively converts one enantiomer in a racemic mixture to another compound, after which the unreacted enantiomer and the new compound are separated. The reaction by lipase in Section 5.10B is an example of this type of resolution. Chromatography using chiral media is also widely used to resolve enantiomers. This approach is applied in high-performance liquid chromatography (HPLC) as well as in other forms of chromatography. Diastereomeric interactions between molecules of the racemic mixture and the chiral chromatography medium cause enantiomers of the racemate to move through the chromatography apparatus at different rates. The enantiomers are then collected separately as they elute from the chromatography device. (See "*The Chemistry of* . . . HPLC Resolution of Enantiomers," Section 20.3.)

5.17 *Compounds with Chirality Centers Other than Carbon*

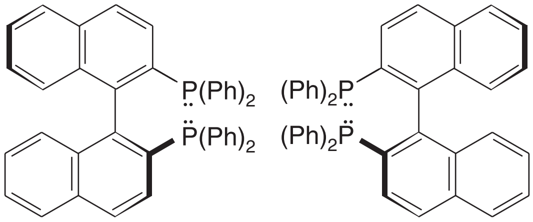
Any tetrahedral atom with four different groups attached to it is a chirality center. Shown here are general formulas of compounds whose molecules contain chirality centers other than carbon. Silicon and germanium are in the same group of the periodic table as carbon. They form tetrahedral compounds as carbon does. When four different groups are situated around the central atom in silicon, germanium, and nitrogen compounds, the molecules are chiral and the enantiomers can, in principle, be separated. Sulfoxides, like certain examples of other functional groups where one of the four groups is a nonbonding electron pair, are also chiral. This is not the case for amines, however (Section 20.2B):



5.18 *Chiral Molecules That Do Not Possess a Chirality Center*

A molecule is chiral if it is not superposable on its mirror image. The presence of a tetrahedral atom with four different groups is only one type of chirality center, however. While most of the chiral molecules we shall encounter have chirality centers, there are other structural attributes that can confer chirality on a molecule. For example, there are compounds that have such large rotational barriers between conformers that individual conformational isomers can be separated and purified, and some of these conformational isomers are stereoisomers.

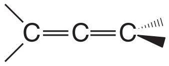
Conformational isomers that are stable, isolable compounds are called **atropisomers**. The existence of chiral atropisomers has been exploited to great effect in the development of chiral catalysts for stereoselective reactions. An example is BINAP, shown below in its enantiomeric forms:



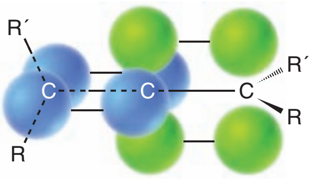
***(S)*-BINAP** ***(R)*-BINAP**

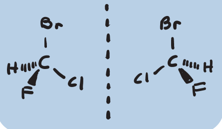
The origin of chirality in BINAP is the restricted rotation about the bond between the two nearly perpendicular naphthalene rings. This torsional barrier leads to two resolvable enantiomeric conformers, (*S*)- and (*R*)-BINAP. When each enantiomer is used as a ligand for metals such as ruthenium and rhodium (bound by unshared electron pairs on the phosphorus atoms), chiral organometallic complexes result that are capable of catalyzing stereoselective hydrogenation and other important industrial reactions. The significance of chiral ligands is highlighted by the industrial synthesis each year of approximately 3500 *tons* of (−)-menthol using an isomerization reaction involving a rhodium (*S*)-BINAP catalyst.

Allenes are compounds that also exhibit stereoisomerism. Allenes are molecules that contain the following double-bond sequence:



The planes of the π bonds of allenes are perpendicular to each other:





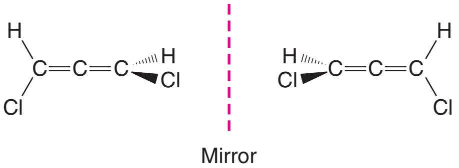


Figure 5.22 Enantiomeric forms of 1,3-dichloroallene. These two molecules are nonsuperposable mirror images of each other and are therefore chiral. They do not possess a tetrahedral atom with four different groups, however.

This geometry of the π bonds causes the groups attached to the end carbon atoms to lie in perpendicular planes, and, because of this, allenes with different substituents on the end carbon atoms are chiral (Fig. 5.22). (Allenes do not show cis–trans isomerism.)

In This Chapter

In this chapter you learned that the handedness of life begins at the molecular level. Molecular recognition, signaling, and chemical reactions in living systems often hinge on the handedness of chiral molecules. Molecules that bear four different groups at a tetrahedral carbon atom are chiral if they are nonsuperposable with their mirror image. The atoms bearing four different groups are called chirality centers.

Mirror planes of symmetry have been very important to our discussion. If we want to draw the enantiomer of a molecule, one way to do so is to draw the molecule as if it were reflected in a mirror. If a mirror plane of symmetry exists *within* a molecule, then it is achiral (not chiral), even if it contains chirality centers. Using mirror planes to test for symmetry is an important technique.

In this chapter you learned how to give unique names to chiral molecules using the Cahn–Ingold–Prelog *R,S*–system. You have also exercised your mind's eye in visualizing molecular structures in three dimensions, and you have refined your skill at drawing three-dimensional molecular formulas. You learned that pairs of enantiomers have identical physical properties except for the equal and opposite rotation of plane-polarized light, whereas diastereomers have different physical properties from one another. Interactions between each enantiomer of a chiral molecule and any other chiral material lead to diastereomeric interactions, which lead to different physical properties that can allow the separation of enantiomers.

Chemistry happens in three dimensions. Now, with the information from this chapter building on fundamentals you have learned about molecular shape and polarity in earlier chapters, you are ready to embark on your study of the reactions of organic molecules. Practice drawing molecules that show three dimensions at chirality centers, practice naming molecules, and label their regions of partial positive and negative charge. Paying attention to these things will help you learn about the reactivity of molecules in succeeding chapters. Most important of all, do your homework!

Key Terms and Concepts

The key terms and concepts that are highlighted in **bold, blue text** within the chapter are defined in the glossary (at the back of the book) and have hyperlinked definitions in the accompanying ***WileyPLUS*** course (www.wileyplus.com).

Problems

Note to Instructors: Many of the homework problems are available for assignment via *WileyPLUS*, an online teaching and learning solution.

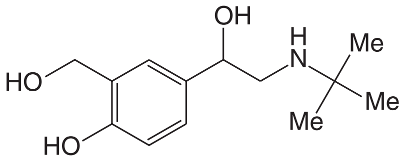
CHIRALITY AND STEREOISOMERISM

5.33 Which of the following are chiral and, therefore, capable of existing as enantiomers?

(a) 1,3-Dichlorobutane
(b) 1,2-Dibromopropane
(c) 1,5-Dichloropentane
(d) 3-Ethylpentane
(e) 2-Bromobicyclo[1.1.0]butane
(f) 2-Fluorobicyclo[2.2.2]octane
(g) 2-Chlorobicyclo[2.1.1]hexane
(h) 5-Chlorobicyclo[2.1.1]hexane

5.34 **(a)** How many carbon atoms does an alkane (not a cycloalkane) need before it is capable of existing in enantiomeric forms? **(b)** Give correct names for two sets of enantiomers with this minimum number of carbon atoms.

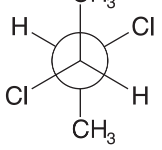
5.35 Albuterol, shown here, is a commonly prescribed asthma medication. For either enantiomer of albuterol, draw a three-dimensional formula using dashes and wedges for bonds that are not in the plane of the paper. Choose a perspective that allows as many carbon atoms as possible to be in the plane of the paper, and show all unshared electron pairs and hydrogen atoms (except those on the methyl groups labeled Me). Specify the (*R,S*) configuration of the enantiomer you drew.



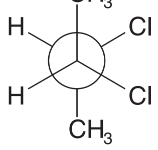
Albuterol

5.36 **(a)** Write the structure of 2,2-dichlorobicyclo[2.2.1]heptane. **(b)** How many chirality centers does it contain? **(c)** How many stereoisomers are predicted by the 2^n rule? **(d)** Only one pair of enantiomers is possible for 2,2-dichlorobicyclo[2.2.1]heptane. Explain.

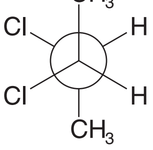
5.37 Shown below are Newman projection formulas for (*R,R*)-, (*S,S*)-, and (*R,S*)-2,3-dichlorobutane. **(a)** Which is which? **(b)** Which formula is a meso compound?



A



B

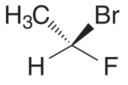


C

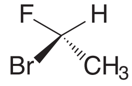
5.38 Write appropriate structural formulas for **(a)** a cyclic molecule that is a constitutional isomer of cyclohexane, **(b)** molecules with the formula C_6H_{12} that contain one ring and that are enantiomers of each other, **(c)** molecules with the formula C_6H_{12} that contain one ring and that are diastereomers of each other, **(d)** molecules with the formula C_6H_{12} that contain no ring and that are enantiomers of each other, and **(e)** molecules with the formula C_6H_{12} that contain no ring and that are diastereomers of each other.

5.39 Consider the following pairs of structures. Designate each chirality center as (*R*) or (*S*) and identify the relationship between them by describing them as representing enantiomers, diastereomers, constitutional isomers, or two molecules of the same compound. Use handheld molecular models to check your answers.

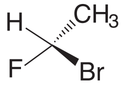
(a)



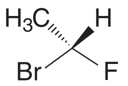
and



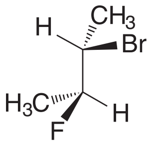
(b)



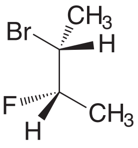
and



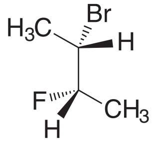
(c)



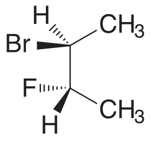
and



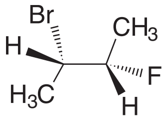
(d)



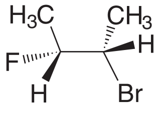
and



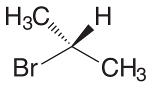
(e)



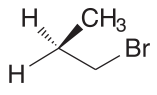
and



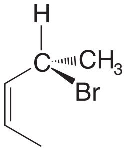
(f)



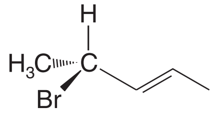
and



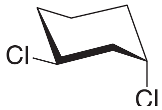
(g)



and

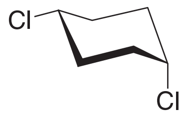


(h)



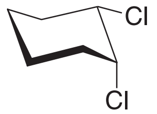
and

(i)



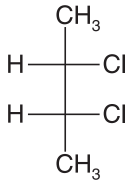
and

(j)

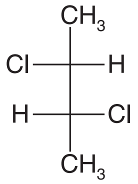


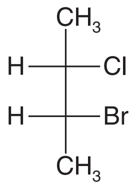
and

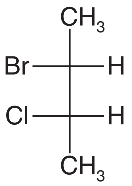
(k)

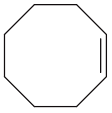
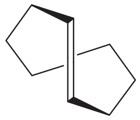


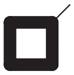
and

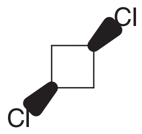


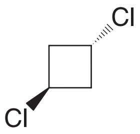
(l) 

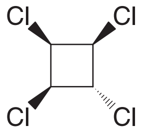
and 

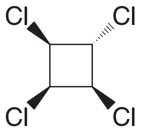
(m)  and 

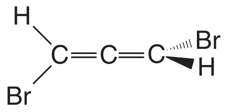
(n)  and 

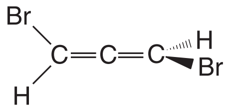
(o) 

and 

(p) 

and 

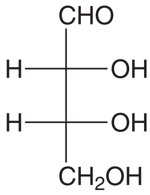
(q) 

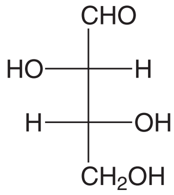
and 

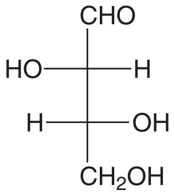
5.40 Discuss the anticipated stereochemistry of each of the following compounds.

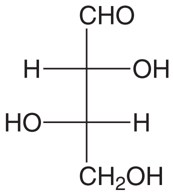
(a) $ClCH{=}C{=}C{=}CHCl$ **(b)** $CH_2{=}C{=}C{=}CHCl$ **(c)** $ClCH{=}C{=}C{=}CCl_2$

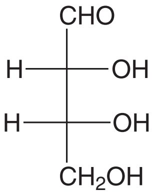
5.41 Tell whether the compounds of each pair are enantiomers, diastereomers, constitutional isomers, or not isomeric.

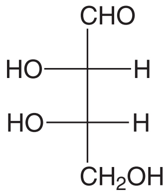
(a) 

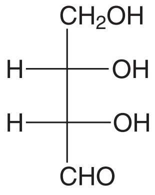
and 

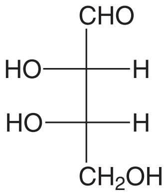
(b) 

and 

(c) 

and 

(d) 

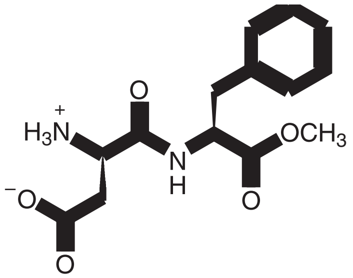
and 

5.42 A compound **D** with the molecular formula C_6H_{12} is optically inactive but can be resolved into enantiomers. On catalytic hydrogenation, **D** is converted to **E** (C_6H_{14}) and **E** is optically inactive. Propose structures for **D** and **E**.

5.43 Compound **F** has the molecular formula C_5H_8 and is optically active. On catalytic hydrogenation **F** yields **G** (C_5H_{12}) and **G** is optically inactive. Propose structures for **F** and **G**.

5.44 Compound **H** is optically active and has the molecular formula C_6H_{10}. On catalytic hydrogenation **H** is converted to **I** (C_6H_{12}) and **I** is optically inactive. Propose structures for **H** and **I**.

5.45 Aspartame is an artificial sweetener. Give the (*R,S*) designation for each chirality center of aspartame.



Aspartame

5.46 There are four dimethylcyclopropane isomers. **(a)** Write three-dimensional formulas for these isomers. **(b)** Which of the isomers are chiral? **(c)** If a mixture consisting of 1 mol of each of these isomers were subjected to simple gas chromatography (an analytical method that can separate compounds according to boiling point), how many fractions would be obtained and which compounds would each fraction contain? **(d)** How many of these fractions would be optically active?

5.47 (Use models to solve this problem.) **(a)** Write a conformational structure for the most stable conformation of *trans*-1,2-diethylcyclohexane and write its mirror image. **(b)** Are these two molecules superposable? **(c)** Are they interconvertible through a ring "flip"? **(d)** Repeat the process in part (a) with *cis*-1,2-diethylcyclohexane. **(e)** Are these structures superposable? **(f)** Are they interconvertible?

5.48 (Use models to solve this problem.) **(a)** Write a conformational structure for the most stable conformation of *trans*-1,4-diethylcyclohexane and for its mirror image. **(b)** Are these structures superposable? **(c)** Do they represent enantiomers? **(d)** Does *trans*-1,4-diethylcyclohexane have a stereoisomer, and if so, what is it? **(e)** Is this stereoisomer chiral?

5.49 (Use models to solve this problem.) Write conformational structures for all of the stereoisomers of 1,3-diethylcyclohexane. Label pairs of enantiomers and meso compounds if they exist.

Challenge Problems

5.50 Tartaric acid [$HO_2CCH(OH)CH(OH)CO_2H$] was an important compound in the history of stereochemistry. Two naturally occurring forms of tartaric acid are optically inactive. One optically inactive form has a melting point of 210–212°C, the other a melting point of 140°C. The inactive tartaric acid with a melting point of 210–212°C can be separated into two optically active forms of tartaric acid with the same melting point (168–170°C). One optically active tartaric acid has $[\alpha]_D^{25} = +12$, and the other, $[\alpha]_D^{25} = -12$. All attempts to separate the other inactive tartaric acid (melting point 140°C) into optically active compounds fail. **(a)** Write the three-dimensional structure of the tartaric acid with melting point 140°C. **(b)** Write structures for the optically active tartaric acids with melting points of 168–170°C. **(c)** Can you determine from the formulas which tartaric acid in (b) has a positive rotation and which has a negative rotation? **(d)** What is the nature of the form of tartaric acid with a melting point of 210–212°C?

5.51 **(a)** An aqueous solution of pure stereoisomer X of concentration 0.10 g mL^{-1} had an observed rotation of $-30°$ in a 1.0-dm tube at 589.6 nm (the sodium D line) and 25°C. What do you calculate its $[\alpha]_D$ to be at this temperature?

(b) Under identical conditions but with concentration 0.050 g mL^{-1}, a solution of X had an observed rotation of $+165°$. Rationalize how this could be and recalculate $[\alpha]_D$ for stereoisomer X.

(c) If the optical rotation of a substance studied at only one concentration is zero, can it definitely be concluded to be achiral? Racemic?

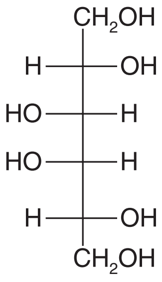
5.52 If a sample of a pure substance that has two or more chirality centers has an observed rotation of zero, it could be a racemate. Could it possibly be a pure stereoisomer? Could it possibly be a pure enantiomer?

5.53 Unknown Y has a molecular formula of $C_3H_6O_2$. It contains one functional group that absorbs infrared radiation in the 3200–3550-cm^{-1} region (when studied as a pure liquid; i.e., "neat"), and it has no absorption in the 1620–1780-cm^{-1} region. No carbon atom in the structure of Y has more than one oxygen atom bonded to it, and Y can exist in two (and only two) stereoisomeric forms. What are the structures of these forms of Y?

Learning Group Problems

1. Streptomycin is an antibiotic that is especially useful against penicillin-resistant bacteria. The structure of streptomycin is shown in Section 22.17. **(a)** Identify all of the chirality centers in the structure of streptomycin. **(b)** Assign the appropriate (*R*) or (*S*) designation for the configuration of each chirality center in streptomycin.

2. D-Galactitol is one of the toxic compounds produced by the disease galactosemia. Accumulation of high levels of D-galactitol causes the formation of cataracts. A Fischer projection for D-galactitol is shown at right:



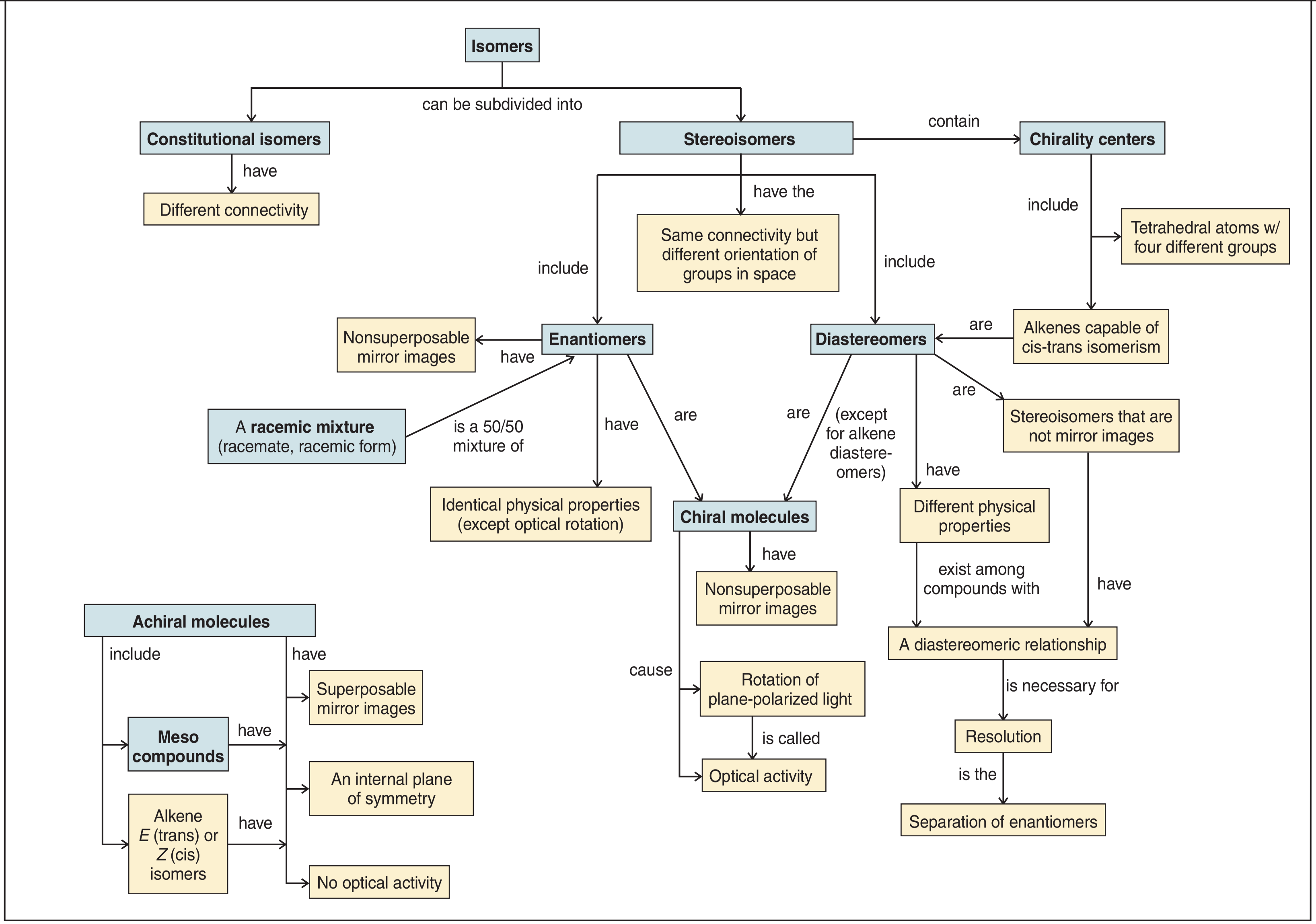
(a) Draw a three-dimensional structure for D-galactitol.

(b) Draw the mirror image of D-galactitol and write its Fischer projection formula.

(c) What is the stereochemical relationship between D-galactitol and its mirror image?

3. Cortisone is a natural steroid that can be isolated from the adrenal cortex. It has anti-inflammatory properties and is used to treat a variety of disorders (e.g., as a topical application for common skin diseases). The structure of cortisone is shown in Section 23.4D. **(a)** Identify all of the chirality centers in cortisone. **(b)** Assign the appropriate (*R*) or (*S*) designation for the configuration of each chirality center in cortisone.

CONCEPT MAP


Isomers
can be subdivided into
Constitutional isomers
have
Different connectivity
Stereoisomers
contain
Chirality centers
include
Tetrahedral atoms w/ four different groups
Alkenes capable of cis-trans isomerism
are
have the
Same connectivity but different orientation of groups in space
include
include
Enantiomers
have
Nonsuperposable mirror images
A racemic mixture (racemate, racemic form)
is a 50/50 mixture of
have
Identical physical properties (except optical rotation)
are
Diastereomers
are
(except for alkene diastereomers)
are
Stereoisomers that are not mirror images
have
Different physical properties
exist among compounds with
have
A diastereomeric relationship
is necessary for
Resolution
is the
Separation of enantiomers
Chiral molecules
have
Nonsuperposable mirror images
cause
Rotation of plane-polarized light
is called
Optical activity
Achiral molecules
include
Meso compounds
Alkene E (trans) or Z (cis) isomers
have
have
have
Superposable mirror images
An internal plane of symmetry
No optical activity