CHAPTER 1

MOLECULAR FORMULAS AND WHAT CAN BE LEARNED FROM THEM

Before attempting to deduce the structure of an unknown organic compound from an examination of its spectra, we can simplify the problem somewhat by examining the molecular formula of the substance. The purpose of this chapter is to describe how the molecular formula of a compound is determined and how structural information may be obtained from that formula. The chapter reviews both the modern and classical *quantitative methods* of determining the molecular formula. While use of the mass spectrometer (Section 1.6 and Chapter 8) can supplant many of these quantitative analytical methods, they are still in use. Many journals still require that a satisfactory quantitative elemental analysis (Section 1.1) be obtained prior to the publication of research results.

1.1 ELEMENTAL ANALYSIS AND CALCULATIONS

The classical procedure for determining the molecular formula of a substance involves three steps:

1. A **qualitative elemental analysis** to find out what types of atoms are present . . . C, H, N, O, S, Cl, and so on.

2. A **quantitative elemental analysis** (or **microanalysis**) to find out the relative numbers (percentages) of each distinct type of atom in the molecule.

3. A **molecular mass** (or **molecular weight**) **determination**.

The first two steps establish an **empirical formula** for the compound. When the results of the third procedure are known, a **molecular formula** is found.

Virtually all organic compounds contain carbon and hydrogen. In most cases, it is not necessary to determine whether these elements are present in a sample: their presence is assumed. However, if it should be necessary to demonstrate that either carbon or hydrogen is present in a compound, that substance may be burned in the presence of excess oxygen. If the combustion produces carbon dioxide, carbon must be present; if combustion produces water, hydrogen atoms must be present. Today, the carbon dioxide and water can be detected by gas chromatographic methods. Sulfur atoms are converted to sulfur dioxide; nitrogen atoms are often chemically reduced to nitrogen gas following their combustion to nitrogen oxides. Oxygen can be detected by the ignition of the compound in an atmosphere of hydrogen gas; the product is water. Currently, all such analyses are performed by gas chromatography, a method that can also determine the relative amounts of each of these gases. If the amount of the original sample is known, it can be entered, and the computer can calculate the **percentage composition** of the sample.

Unless you work in a large company or in one of the larger universities, it is quite rare to find a research laboratory in which elemental analyses are performed on site. It requires too much time to set up the apparatus and keep it operating within the limits of suitable accuracy and precision. Usually, samples are sent to a commercial **microanalytical laboratory** that is prepared to do this work routinely and to vouch for the accuracy of the results.

1

Before the advent of modern instrumentation, the combustion of the precisely weighed sample was carried out in a cylindrical glass tube, contained within a furnace. A stream of oxygen was passed through the heated tube on its way to two other sequential, unheated tubes that contained chemical substances that would absorb first the water ($MgClO_4$) and then the carbon dioxide (NaOH/silica). These preweighed absorption tubes were detachable and were removed and reweighed to determine the amounts of water and carbon dioxide formed. The percentages of carbon and hydrogen in the original sample were calculated by simple stoichiometry. Table 1.1 shows a sample calculation.

Notice in this calculation that the amount of oxygen was determined by difference, a common practice. In a sample containing only C, H, and O, one needs to determine the percentages of only C and H; oxygen is assumed to be the unaccounted-for portion. You may also apply this practice in situations involving elements other than oxygen; if all but one of the elements is determined, the last one can be determined by difference. Today, most calculations are carried out automatically by the computerized instrumentation. Nevertheless, it is often useful for a chemist to understand the fundamental principles of the calculations.

Table 1.2 shows how to determine the **empirical formula** of a compound from the percentage compositions determined in an analysis. Remember that the empirical formula expresses the simplest whole-number ratios of the elements and may need to be multiplied by an integer to obtain the true **molecular formula.** To determine the value of the multiplier, a molecular mass is required. Determination of the molecular mass is discussed in the next section.

For a totally unknown compound (unknown chemical source or history) you will have to use this type of calculation to obtain the suspected empirical formula. However, if you have prepared the compound from a known precursor by a well-known reaction, you will have an idea of the structure of the compound. In this case, you will have calculated the expected percentage composition of your

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TABLE 1.1
CALCULATION OF PERCENTAGE COMPOSITION
FROM COMBUSTION DATA

$C_xH_yO_z$ + excess $O_2 \longrightarrow$ $x\,CO_2$ $+ y/2\,H_2O$
9.83 mg   23.26 mg   9.52 mg

$$\text{millimoles } CO_2 = \frac{23.26 \text{ mg } CO_2}{44.01 \text{ mg/mmole}} = 0.5285 \text{ mmoles } CO_2$$

mmoles CO_2 = mmoles C in original sample

(0.5285 mmoles C)(12.01 mg/mmole C) = 6.35 mg C in original sample

$$\text{millimoles } H_2O = \frac{9.52 \text{ mg } H_2O}{18.02 \text{ mg/mmole}} = 0.528 \text{ mmoles } H_2O$$

$$(0.528 \text{ mmoles } H_2O)\left(\frac{2 \text{ mmoles H}}{1 \text{ mmole } H_2O}\right) = 1.056 \text{ mmoles H in original sample}$$

(1.056 mmoles H)(1.008 mg/mmole H) = 1.06 mg H in original sample

$$\% \, C = \frac{6.35 \text{ mg C}}{9.83 \text{ mg sample}} \times 100 = 64.6\%$$

$$\% \, H = \frac{1.06 \text{ mg H}}{9.83 \text{ mg sample}} \times 100 = 10.8\%$$

% O = 100 − (64.6 + 10.8) = 24.6%

T A B L E 1 . 2
CALCULATION OF EMPIRICAL FORMULA

Using a 100-g sample:

   64.6% of C = 64.6 g

   10.8% of H = 10.8 g

   24.6% of O = $\dfrac{24.6 \text{ g}}{100.0 \text{ g}}$

   moles C = $\dfrac{64.6 \text{ g}}{12.01 \text{ g/mole}}$ = 5.38 moles C

   moles H = $\dfrac{10.8 \text{ g}}{1.008 \text{ g/mole}}$ = 10.7 moles H

   moles O = $\dfrac{24.6 \text{ g}}{16.0 \text{ g/mole}}$ = 1.54 moles O

giving the result

   $C_{5.38}H_{10.7}O_{1.54}$

Converting to the simplest ratio:

   $C_{\frac{5.38}{1.54}} H_{\frac{10.7}{1.54}} O_{\frac{1.54}{1.54}} = C_{3.49}H_{6.95}O_{1.00}$

which approximates

   $C_{3.50}H_{7.00}O_{1.00}$

or

   $C_{7}H_{14}O_{2}$

sample in advance (from its postulated structure) and will use the analysis to verify your hypothesis. When you perform these calculations, be sure to use the full molecular weights as given in the periodic chart and do not round off until you have completed the calculation. The final result should be good to two decimal places: four significant figures if the percentage is between 10 and 100; three figures if it is between 0 and 10. If the analytical results do not agree with the calculation, the sample may be impure, or you may have to calculate a new empirical formula to discover the identity of the unexpected structure. To be accepted for publication, most journals require the percentages found to be *less than 0.4% off from the calculated value*. Most microanalytical laboratories can easily obtain accuracy well below this limit provided the sample is pure.

   In Figure 1.1, a typical situation for the use of an analysis in research is shown. Professor Amyl Carbon, or one of his students, prepared a compound believed to be the epoxynitrile with the structure shown at the bottom of the first form. A sample of this liquid compound (25 μL) was placed in a small vial correctly labeled with the name of the submitter and an identifying code (usually one that corresponds to an entry in the research notebook). Only a small amount of the sample is required, usually a few milligrams of a solid or a few microliters of a liquid. A Request for Analysis form must be filled out and submitted along with the sample. The sample form on the left side of the figure shows the type of information that must be submitted. In this case, the professor calculated the expected results for C, H, and N and the expected formula and molecular weight. Note that the compound also contains oxygen, but that there was no request for an oxygen analysis. Two other samples were also submitted at the same time. After a short time, typically within a week, the

Microanalytical Company, Inc.

REQUEST FOR ANALYSIS FORM

Date: _October 30, 2006_

Report To: _Professor Amyl Carbon_
Department of Chemistry
Western Washington University
Bellingham, WA 98225

Sample No: _PAC599A_ P.O. No : _PO 2349_

Report By: AirMail Phone (Email)
(circle one)   _pac@www.edu_

Elements to Analyze: _C, H, N_

Other Elements Present : _O_

[X] Single Analysis [ ] Duplicate Analysis

[ ] Duplicate only if results are not in range

M.P. _____ B.P. _69_ ˚C @ _2.3_ mmHg
Sensitive to : _____ Weigh under N? Y (N)
Dry the Sample? Y (N) Details: _____

[ ] Hygroscopic [ ] Volatile [ ] Explosive

THEORY OR RANGE

%C _67.17_ Amount Provided _25 μL_
%H _8.86_   Stucture:
%N _11.19_
%O _____
%Other _____ Comments: _$C_7H_{11}NO$_
Mol. Wt. _125.17_

Microanalytical Company, Inc.

November 25, 2006

Professor Amyl Carbon
Department of Chemistry
Western Washington University
Bellingham, WA

RESULTS OF ANALYSIS

Sample ID	Carbon (%)	Hydrogen (%)	Nitrogen (%)
PAC599A	67.39	9.22	11.25
PAC589B	64.98	9.86	8.03
PAC603	73.77	8.20	–

Dr. B. Grant Poohbah,
_Ph.D._

Director of Analytical Services
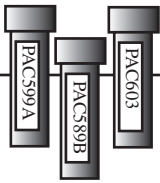Microanalytical Company, Inc

FIGURE 1.1 Sample microanalysis forms. Shown on the left is a typical submission form that is sent with the samples. (The three shown here in labeled vials were all sent at the same time.) Each sample needs its own form. In the background on the right is the formal letter that reported the results. Were the results obtained for sample PAC599A satisfactory?

results were reported to Professor Carbon as an email (see the request on the form). At a later date, a formal letter (shown in the background on the right-hand side) is sent to verify and authenticate the results. Compare the values in the report to those calculated by Professor Carbon. Are they within the accepted range? If not, the analysis will have to be repeated with a freshly purified sample, or a new possible structure will have to be considered.

  Keep in mind that in an actual laboratory situation, when you are trying to determine the molecular formula of a totally new or previously unknown compound, you will have to allow for some variance in the quantitative elemental analysis. Other data can help you in this situation since infrared (Chapter Two) and nuclear magnetic resonance (NMR) (Chapter Three) data will also suggest a possible structure or at least some of its prominent features. Many times, these other data will be less sensitive to small amounts of impurities than the microanalysis.

1.2 DETERMINATION OF MOLECULAR MASS

The next step in determining the molecular formula of a substance is to determine the weight of one mole of that substance. This may be accomplished in a variety of ways. Without knowledge of the molecular mass of the unknown, there is no way of determining whether the empirical formula, which is determined directly from elemental analysis, is the true formula of the substance or whether the empirical formula must be multiplied by some integral factor to obtain the molecular formula. In the example cited in Section 1.1, without knowledge of the molecular mass of the unknown, it is impossible to tell whether the molecular formula is $C_7H_{14}O_2$ or $C_{14}H_{28}O_4$.

In a modern laboratory, the molecular mass is determined using mass spectrometry. The details of this method and the means of determining molecular mass can be found in Section 1.6 and Chapter 8, Section 8.6. This section reviews some classical methods of obtaining the same information.

An old method that is used occasionally is the **vapor density method.** In this method, a known volume of gas is weighed at a known temperature. After converting the volume of the gas to standard temperature and pressure, we can determine what fraction of a mole that volume represents. From that fraction, we can easily calculate the molecular mass of the substance.

Another method of determining the molecular mass of a substance is to measure the freezing-point depression of a solvent that is brought about when a known quantity of test substance is added. This is known as a **cryoscopic method.** Another method, which is used occasionally, is **vapor pressure osmometry,** in which the molecular weight of a substance is determined through an examination of the change in vapor pressure of a solvent when a test substance is dissolved in it.

If the unknown substance is a carboxylic acid, it may be titrated with a standardized solution of sodium hydroxide. By use of this procedure, a **neutralization equivalent** can be determined. The neutralization equivalent is identical to the equivalent weight of the acid. If the acid has only one carboxyl group, the neutralization equivalent and the molecular mass are identical. If the acid has more than one carboxyl group, the neutralization equivalent is equal to the molecular mass of the acid divided by the number of carboxyl groups. Many phenols, especially those substituted by electron-withdrawing groups, are sufficiently acidic to be titrated by this same method, as are sulfonic acids.

1.3 MOLECULAR FORMULAS

Once the molecular mass and the empirical formula are known, we may proceed directly to the **molecular formula.** Often, the empirical formula weight and the molecular mass are the same. In such cases, the empirical formula is also the molecular formula. However, in many cases, the empirical formula weight is less than the molecular mass, and it is necessary to determine how many times the empirical formula weight can be divided into the molecular mass. The factor determined in this manner is the one by which the empirical formula must be multiplied to obtain the molecular formula.

Ethane provides a simple example. After quantitative element analysis, the empirical formula for ethane is found to be **CH_3.** A molecular mass of 30 is determined. The empirical formula weight of ethane, 15, is half of the molecular mass, 30. Therefore, the molecular formula of ethane must be $2(CH_3)$ or C_2H_6.

For the sample unknown introduced earlier in this chapter, the empirical formula was found to be $C_7H_{14}O_2$. The formula weight is 130. If we assume that the molecular mass of this substance was determined to be 130, we may conclude that the empirical formula and the molecular formula are identical, and that the molecular formula must be $C_7H_{14}O_2$.

1.4 INDEX OF HYDROGEN DEFICIENCY

Frequently, a great deal can be learned about an unknown substance simply from knowledge of its molecular formula. This information is based on the following general molecular formulas:

alkane	C_nH_{2n+2}	} Difference of 2 hydrogens
cycloalkane or alkene	C_nH_{2n}	
alkyne	C_nH_{2n-2}	} Difference of 2 hydrogens

Notice that each time a ring or π bond is introduced into a molecule, the number of hydrogens in the molecular formula is reduced by *two*. For every *triple bond* (*two* π bonds), the molecular formula is reduced by four. This is illustrated in Figure 1.2.

When the molecular formula for a compound contains noncarbon or nonhydrogen elements, the ratio of carbon to hydrogen may change. Following are three simple rules that may be used to predict how this ratio will change:

1. To convert the formula of an open-chain, saturated hydrocarbon to a formula containing Group V elements (N, P, As, Sb, Bi), one additional hydrogen atom must be *added* to the molecular formula for each such Group V element present. In the following examples, each formula is correct for a two-carbon acyclic, saturated compound:

$$C_2H_6, \qquad C_2H_7N, \qquad C_2H_8N_2, \qquad C_2H_9N_3$$

2. To convert the formula of an open-chain, saturated hydrocarbon to a formula containing Group VI elements (O, S, Se, Te), *no change* in the number of hydrogens is required. In the following examples, each formula is correct for a two-carbon, acyclic, saturated compound:
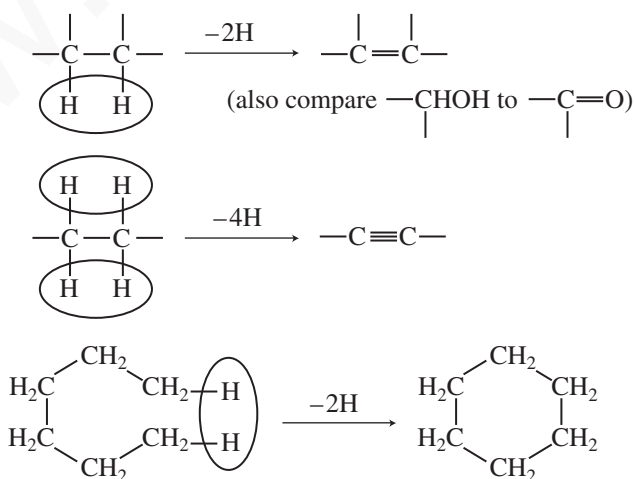
$$C_2H_6, \qquad C_2H_6O, \qquad C_2H_6O_2, \qquad C_2H_6O_3$$

FIGURE 1.2 Formation of rings and double bonds. Formation of each ring or double bond causes the loss of 2H.

3. To convert the formula of an open-chain, saturated hydrocarbon to a formula containing Group VII elements (F, Cl, Br, I), one hydrogen must be *subtracted* from the molecular formula for each such Group VII element present. In the following examples, each formula is correct for a two-carbon, acyclic, saturated compound:

$$C_2H_6, \qquad C_2H_5F, \qquad C_2H_4F_2, \qquad C_2H_3F_3$$

Table 1.3 presents some examples that should demonstrate how these correction numbers were determined for each of the heteroatom groups.

The **index of hydrogen deficiency** (sometimes called the **unsaturation index**) is the number of π bonds and/or rings a molecule contains. It is determined from an examination of the molecular formula of an unknown substance and from a comparison of that formula with a formula for a corresponding acyclic, saturated compound. The difference in the number of hydrogens between these formulas, when divided by 2, gives the index of hydrogen deficiency.

The index of hydrogen deficiency can be very useful in structure determination problems. A great deal of information can be obtained about a molecule before a single spectrum is examined. For example, a compound with an index of **one** must have one double bond or one ring, but it cannot have both structural features. A quick examination of the infrared spectrum could confirm the presence of a double bond. If there were no double bond, the substance would have to be cyclic and saturated. A compound with an index of **two** could have a triple bond, or it could have two double bonds, two rings, or one of each. Knowing the index of hydrogen deficiency of a substance, the chemist can proceed directly to the appropriate regions of the spectra to confirm the presence or absence of π bonds or rings. Benzene contains one ring and three "double bonds" and thus has an index of hydrogen deficiency of **four.** Any substance with an index of *four* or more may contain a benzenoid ring; a substance with an index less than *four* cannot contain such a ring.

To determine the index of hydrogen deficiency for a compound, apply the following steps:

1. Determine the formula for the saturated, acyclic hydrocarbon containing the same number of carbon atoms as the unknown substance.

2. Correct this formula for the nonhydrocarbon elements present in the unknown. Add one hydrogen atom for each Group V element present and subtract one hydrogen atom for each Group VII element present.

3. Compare this formula with the molecular formula of the unknown. Determine the number of hydrogens by which the two formulas differ.

4. Divide the difference in the number of hydrogens by **two** to obtain the index of hydrogen deficiency. This equals the number of π bonds and/or rings in the structural formula of the unknown substance.

TABLE 1.3
CORRECTIONS TO THE NUMBER OF HYDROGEN ATOMS WHEN GROUP V AND VII HETEROATOMS ARE INTRODUCED (GROUP VI HETEROATOMS DO NOT REQUIRE A CORRECTION)

| Group	Example	Correction	Net Change
V	C—H → C—NH$_2$	+1	Add nitrogen, add 1 hydrogen
VI	C—H → C—OH	0	Add oxygen (no hydrogen)
VII	C—H → C—Cl	−1	Add chlorine, lose 1 hydrogen

The following examples illustrate how the index of hydrogen deficiency is determined and how that information can be applied to the determination of a structure for an unknown substance.

■ EXAMPLE 1

The unknown substance introduced at the beginning of this chapter has the molecular formula $C_7H_{14}O_2$.

1. Using the general formula for a saturated, acyclic hydrocarbon (C_nH_{2n+2}, where $n = 7$), calculate the formula C_7H_{16}.

2. Correction for oxygens (no change in the number of hydrogens) gives the formula $C_7H_{16}O_2$.

3. The latter formula differs from that of the unknown by two hydrogens.

4. The index of hydrogen deficiency equals **one.** There must be one ring or one double bond in the unknown substance.

Having this information, the chemist can proceed immediately to the double-bond regions of the infrared spectrum. There, she finds evidence for a carbon–oxygen double bond (carbonyl group). At this point, the number of possible isomers that might include the unknown has been narrowed considerably. Further analysis of the spectral evidence leads to an identification of the unknown substance as **isopentyl acetate.**

$$CH_3-\overset{\overset{\displaystyle O}{\|}}{C}-O-CH_2-CH_2-\underset{\underset{\displaystyle CH_3}{|}}{CH}-CH_3$$
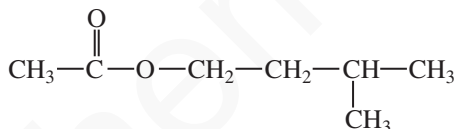
■ EXAMPLE 2

Nicotine has the molecular formula $C_{10}H_{14}N_2$.

1. The formula for a 10-carbon, saturated, acyclic hydrocarbon is $C_{10}H_{22}$.

2. Correction for the two nitrogens (add two hydrogens) gives the formula $C_{10}H_{24}N_2$.

3. The latter formula differs from that of nicotine by 10 hydrogens.

4. The index of hydrogen deficiency equals **five.** There must be some combination of five π bonds and/or rings in the molecule. Since the index is greater than *four,* a benzenoid ring could be included in the molecule.

Analysis of the spectrum quickly shows that a benzenoid ring is indeed present in nicotine. The spectral results indicate no other double bonds, suggesting that another ring, this one saturated, must be present in the molecule. More careful refinement of the spectral analysis leads to a structural formula for nicotine:
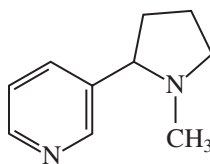
■ **EXAMPLE 3**

Chloral hydrate ("knockout drops") is found to have the molecular formula $C_2H_3Cl_3O_2$.

1. The formula for a two-carbon, saturated, acyclic hydrocarbon is C_2H_6.

2. Correction for oxygens (no additional hydrogens) gives the formula $C_2H_6O_2$.

3. Correction for chlorines (subtract three hydrogens) gives the formula $C_2H_3Cl_3O_2$.

4. This formula and the formula of chloral hydrate correspond exactly.

5. The index of hydrogen deficiency equals **zero.** Chloral hydrate cannot contain rings or double bonds.

Examination of the spectral results is limited to regions that correspond to singly bonded structural features. The correct structural formula for chloral hydrate follows. You can see that all of the bonds in the molecule are single bonds.

$$Cl_3C-\underset{\underset{OH}{|}}{\overset{\overset{OH}{|}}{C}}-H$$
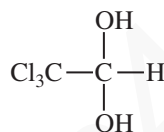
■ **1.5 THE RULE OF THIRTEEN**

High-resolution mass spectrometry provides molecular mass information from which the user can determine the exact molecular formula directly. The discussion on exact mass determination in Chapter 8 explains this process in detail. When such molar mass information is not available, however, it is often useful to be able to generate all the possible molecular formulas for a given mass. By applying other types of spectroscopic information, it may then be possible to distinguish among these possible formulas. A useful method for generating possible molecular formulas for a given molecular mass is the **Rule of Thirteen.**[1]

As a first step in the Rule of Thirteen, we generate a **base formula,** which contains only carbon and hydrogen. The base formula is found by dividing the molecular mass M by 13 (the mass of one carbon plus one hydrogen). This calculation provides a numerator n and a remainder r.

$$\frac{M}{13} = n + \frac{r}{13}$$

The base formula thus becomes

$$C_nH_{n+r}$$

which is a combination of carbons and hydrogens that has the desired molecular mass M.

The **index of hydrogen deficiency** (unsaturation index) U that corresponds to the preceding formula is calculated easily by applying the relationship

$$U = \frac{(n - r + 2)}{2}$$

[1] Bright, J. W., and E. C. M. Chen, "Mass Spectral Interpretation Using the 'Rule of 13,'" *Journal of Chemical Education, 60* (1983): 557.

Of course, you can also calculate the index of hydrogen deficiency using the method shown in Section 1.4.

If we wish to derive a molecular formula that includes other atoms besides carbon and hydrogen, then we must subtract the mass of a combination of carbons and hydrogens that equals the masses of the other atoms being included in the formula. For example, if we wish to convert the base formula to a new formula containing one oxygen atom, then we subtract one carbon and four hydrogens at the same time that we add one oxygen atom. Both changes involve a molecular mass equivalent of 16 ($O = CH_4 = 16$). Table 1.4 includes a number of C/H equivalents for replacement of carbon and hydrogen in the base formula by the most common elements likely to occur in an organic compound.[2]

To comprehend how the Rule of Thirteen might be applied, consider an unknown substance with a molecular mass of 94 amu. Application of the formula provides

$$\frac{94}{13} = 7 + \frac{3}{13}$$

According to the formula, $n = 7$ and $r = 3$. The base formula must be
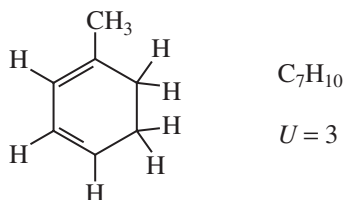
$$C_7H_{10}$$

The index of hydrogen deficiency is

$$U = \frac{(7 - 3 + 2)}{2} = 3$$

TABLE 1.4
CARBON/HYDROGEN EQUIVALENTS FOR SOME COMMON ELEMENTS

Add Element	Subtract Equivalent	Add ΔU	Add Element	Subtract Equivalent	Add ΔU
C	H_{12}	7	^{35}Cl	C_2H_{11}	3
H_{12}	C	-7	^{79}Br	C_6H_7	-3
O	CH_4	1	^{79}Br	C_5H_{19}	4
O_2	C_2H_8	2	F	CH_7	2
O_3	C_3H_{12}	3	Si	C_2H_4	1
N	CH_2	$\frac{1}{2}$	P	C_2H_7	2
N_2	C_2H_4	1	I	C_9H_{19}	0
S	C_2H_8	2	I	$C_{10}H_7$	7

[2] In Table 1.4, the equivalents for chlorine and bromine are determined assuming that the isotopes are ^{35}Cl and ^{79}Br, respectively. Always use this assumption when applying this method.

A substance that fits this formula must contain some combination of three rings or multiple bonds. A possible structure might be


$$C_7H_{10}$$

$$U = 3$$

If we were interested in a substance that had the same molecular mass but that contained one oxygen atom, the molecular formula would become C_6H_6O. This formula is determined according to the following scheme:

1. Base formula $= C_7H_{10}$ $U = 3$
2. Add: $+ O$
3. Subtract: $- CH_4$
4. Change the value of U: $\Delta U = 1$
5. New formula $= C_6H_6O$
6. New index of hydrogen deficiency: $U = 4$

A possible substance that fits these data is

$$C_6H_6O$$

$$U = 4$$
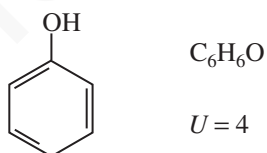
There are additional possible molecular formulas that conform to a molecular mass of 94 amu:

$$C_5H_2O_2 \quad U = 5 \qquad\qquad C_5H_2S \quad U = 5$$
$$C_6H_8N \quad U = 3\tfrac{1}{2} \qquad\qquad CH_3Br \quad U = 0$$
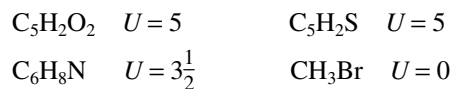
As the formula C_6H_8N shows, any formula that contains an even number of hydrogen atoms but an odd number of nitrogen atoms leads to a fractional value of U, an unlikely choice.

Any compound with a value of U less than zero (i.e., negative) is an impossible combination. Such a value is often an indicator that an oxygen or nitrogen atom must be present in the molecular formula.

When we calculate formulas using this method, if there are not enough hydrogens, we can subtract 1 carbon and add 12 hydrogens (and make the appropriate correction in U). This procedure works only if we obtain a positive value of U. Alternatively we can obtain another potential molecular formula by adding 1 carbon and subtracting 12 hydrogens (and correcting U).

1.6 A QUICK LOOK AHEAD TO SIMPLE USES OF MASS SPECTRA

Chapter 8 contains a detailed discussion of the technique of **mass spectrometry.** See Sections 8.1–8.7 for applications of mass spectrometry to the problems of molecular formula determination. Briefly, the mass spectrometer is an instrument that subjects molecules to a high-energy beam of electrons. This beam of electrons converts molecules to positive ions by removing an electron. The stream of positively charged ions is accelerated along a curved path in a magnetic field. The radius of curvature of the path described by the ions depends on the ratio of the mass of the ion to its charge (the m/z ratio). The ions strike a detector at positions that are determined by the radius of curvature of their paths. The number of ions with a particular mass-to-charge ratio is plotted as a function of that ratio.

The particle with the largest mass-to-charge ratio, assuming that the charge is 1, is the particle that represents the intact molecule with only one electron removed. This particle, called the **molecular ion** (see Chapter 8, Section 8.5), can be identified in the mass spectrum. From its position in the spectrum, its weight can be determined. Since the mass of the dislodged electron is so small, the mass of the molecular ion is essentially equal to the molecular mass of the original molecule. Thus, the mass spectrometer is an instrument capable of providing molecular mass information.

Virtually every element exists in nature in several isotopic forms. The natural abundance of each of these isotopes is known. Besides giving the mass of the molecular ion when each atom in the molecule is the most common isotope, the mass spectrum also gives peaks that correspond to that same molecule with heavier isotopes. The ratio of the intensity of the molecular ion peak to the intensities of the peaks corresponding to the heavier isotopes is determined by the natural abundance of each isotope. Because each type of molecule has a unique combination of atoms, and because each type of atom and its isotopes exist in a unique ratio in nature, the ratio of the intensity of the molecular ion peak to the intensities of the isotopic peaks can provide information about the number of each type of atom present in the molecule.

For example, the presence of bromine can be determined easily because bromine causes a pattern of molecular ion peaks and isotope peaks that is easily identified. If we identify the mass of the molecular ion peak as M and the mass of the isotope peak that is two mass units heavier than the molecular ion as $M + 2$, then the ratio of the intensities of the M and $M + 2$ peaks will be approximately *one to one* when bromine is present (see Chapter 8, Section 8.5, for more details). When chlorine is present, the ratio of the intensities of the M and $M + 2$ peaks will be approximately *three to one.* These ratios reflect the natural abundances of the common isotopes of these elements. Thus, **isotope ratio studies** in mass spectrometry can be used to determine the molecular formula of a substance.

Another fact that can be used in determining the molecular formula is expressed as the **Nitrogen Rule.** This rule states that when the number of nitrogen atoms present in the molecule is odd, the molecular mass will be an odd number; when the number of nitrogen atoms present in the molecule

TABLE 1.5
PRECISE MASSES FOR SUBSTANCES
OF MOLECULAR MASS EQUAL TO 44 amu

Compound	Exact Mass (amu)
CO_2	43.9898
N_2O	44.0011
C_2H_4O	44.0262
C_3H_8	44.0626

is even (or zero), the molecular mass will be an even number. The Nitrogen Rule is explained further in Chapter 8, Section 8.6.

Since the advent of high-resolution mass spectrometers, it is also possible to use very precise mass determinations of molecular ion peaks to determine molecular formulas. When the atomic weights of the elements are determined very precisely, it is found that they do not have exactly integral values. Every isotopic mass is characterized by a small "mass defect," which is the amount by which the mass of the isotope differs from a perfectly integral mass number. The mass defect for every isotope of every element is unique. As a result, a precise mass determination can be used to determine the molecular formula of the sample substance, since every combination of atomic weights at a given nominal mass value will be unique when mass defects are considered. For example, each of the substances shown in Table 1.5 has a nominal mass of 44 amu. As can be seen from the table, their *exact masses,* obtained by adding exact atomic masses, are substantially different when measured to four decimal places.

PROBLEMS

*1. Researchers used a combustion method to analyze a compound used as an antiknock additive in gasoline. A 9.394-mg sample of the compound yielded 31.154 mg of carbon dioxide and 7.977 mg of water in the combustion.

   (a) Calculate the percentage composition of the compound.

   (b) Determine its empirical formula.

*2. The combustion of an 8.23-mg sample of unknown substance gave 9.62 mg CO_2 and 3.94 mg H_2O. Another sample, weighing 5.32 mg, gave 13.49 mg AgCl in a halogen analysis. Determine the percentage composition and empirical formula for this organic compound.

*3. An important amino acid has the percentage composition C 32.00%, H 6.71%, and N 18.66%. Calculate the empirical formula of this substance.

*4. A compound known to be a pain reliever had the empirical formula $C_9H_8O_4$. When a mixture of 5.02 mg of the unknown and 50.37 mg of camphor was prepared, the melting point of a portion of this mixture was determined. The observed melting point of the mixture was 156°C. What is the molecular mass of this substance?

*5. An unknown acid was titrated with 23.1 mL of 0.1 N sodium hydroxide. The weight of the acid was 120.8 mg. What is the equivalent weight of the acid?

*6. Determine the index of hydrogen deficiency for each of the following compounds:

   (a) C_8H_7NO (d) $C_5H_3ClN_4$

   (b) $C_3H_7NO_3$ (e) $C_{21}H_{22}N_2O_2$

   (c) $C_4H_4BrNO_2$

*7. A substance has the molecular formula C_4H_9N. Is there any likelihood that this material contains a triple bond? Explain your reasoning.

*8. (a) A researcher analyzed an unknown solid, extracted from the bark of spruce trees, to determine its percentage composition. An 11.32-mg sample was burned in a combustion apparatus. The carbon dioxide (24.87 mg) and water (5.82 mg) were collected and weighed. From the results of this analysis, calculate the percentage composition of the unknown solid.

   (b) Determine the empirical formula of the unknown solid.

(c) Through mass spectrometry, the molecular mass was found to be 420 g/mole. What is the molecular formula?

(d) How many aromatic rings could this compound contain?

*9. Calculate the molecular formulas for possible compounds with molecular masses of 136; use the Rule of Thirteen. You may assume that the only other atoms present in each molecule are carbon and hydrogen.

(a) A compound with two oxygen atoms

(b) A compound with two nitrogen atoms

(c) A compound with two nitrogen atoms and one oxygen atom

(d) A compound with five carbon atoms and four oxygen atoms

*10. An alkaloid was isolated from a common household beverage. The unknown alkaloid proved to have a molecular mass of 194. Using the Rule of Thirteen, determine a molecular formula and an index of hydrogen deficiency for the unknown. Alkaloids are naturally occurring organic substances that contain **nitrogen.** (*Hint:* There are four nitrogen atoms and two oxygen atoms in the molecular formula. The unknown is **caffeine.** Look up the structure of this substance in *The Merck Index* and confirm its molecular formula.)

*11. The Drug Enforcement Agency (DEA) confiscated a hallucinogenic substance during a drug raid. When the DEA chemists subjected the unknown hallucinogen to chemical analysis, they found that the substance had a molecular mass of 314. Elemental analysis revealed the presence of carbon and hydrogen only. Using the Rule of Thirteen, determine a molecular formula and an index of hydrogen deficiency for this substance. (*Hint:* The molecular formula of the unknown also contains two oxygen atoms. The unknown is **tetrahydrocannabinol,** the active constituent of marijuana. Look up the structure of tetrahydrocannabinol in *The Merck Index* and confirm its molecular formula.)

12. A carbohydrate was isolated from a sample of cow's milk. The substance was found to have a molecular mass of 342. The unknown carbohydrate can be hydrolyzed to form two isomeric compounds, each with a molecular mass of 180. Using the Rule of Thirteen, determine a molecular formula and an index of hydrogen deficiency for the unknown and for the hydrolysis products. (*Hint:* Begin by solving the molecular formula for the 180-amu hydrolysis products. These products have one oxygen atom for every carbon atom in the molecular formula. The unknown is **lactose.** Look up its structure in *The Merck Index* and confirm its molecular formula.)

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*Answers are provided in the chapter, *Answers to Selected Problems*

REFERENCES

O'Neil, M. J., et al., eds. *The Merck Index,* 14th ed., Whitehouse Station, NJ: Merck & Co., 2006.

Pavia, D. L., G. M. Lampman, G. S. Kriz, and R. G. Engel, *Introduction to Organic Laboratory Techniques: A Small Scale Approach,* 2nd ed., Belmont, CA: Brooks-Cole Thomson, 2005.

Pavia, D. L., G. M. Lampman, G. S. Kriz, and R. G. Engel, *Introduction to Organic Laboratory Techniques: A Micro-* *scale Approach,* 4th ed., Belmont, CA: Brooks-Cole Thomson, 2007.

Shriner, R. L., C. K. F. Hermann, T. C. Morrill, D. Y. Curtin, and R. C. Fuson, *The Systematic Identification of Organic Compounds,* 8th ed., New York, NY: John Wiley, 2004.