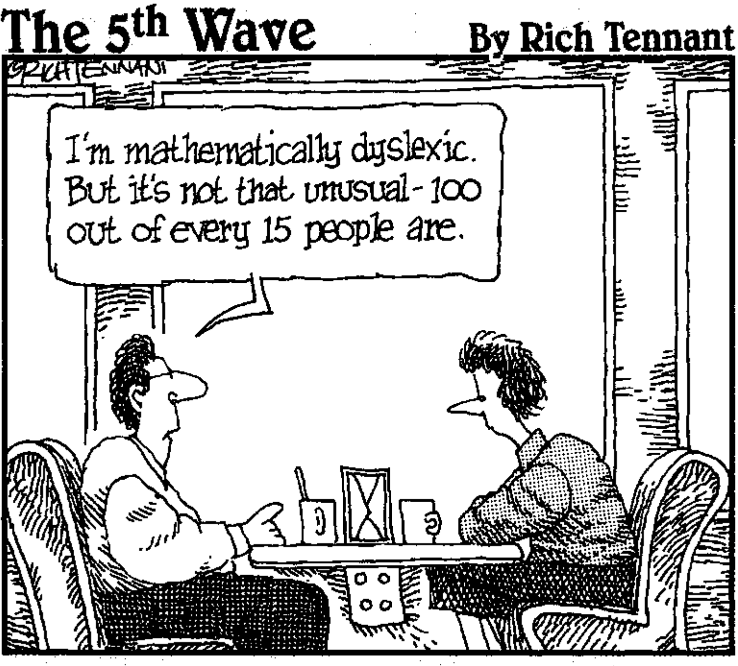
Part I

Basic Concepts of Chemistry



In this part . . .

If you are new to chemistry, it may seem a little frightening. I see students every day who've psyched themselves out by saying so often that they can't do chemistry.

Anyone can figure out chemistry. Anyone can *do* chemistry. If you cook, clean, or simply exist, you're part of the chemical world.

I work with a lot of elementary school children, and they love science. I show them chemical reactions (vinegar plus baking soda, for example), and they go wild. And that's what I hope happens to you.

The chapters of Part I give you a background in chemistry basics. I tell you about matter and the states it can exist in. I talk a little about energy, including the different types and how it's measured. I discuss the microscopic world of the atom and its basic parts. I explain the periodic table, the most useful tool for a chemist. And I cover radioactivity, nuclear reactors, and bombs.

This part takes you on a fun ride, so get your motor running!

Chapter 1

What Is Chemistry, and Why Do I Need to Know Some?

In This Chapter

- Defining the science of chemistry
- Checking out the general areas of chemistry
- Discovering how chemistry is all around you

If you're taking a course in chemistry, you may want to skip this chapter and go right to the area you're having trouble with. But if you bought this book to help you decide whether to take a course in chemistry or to have fun discovering something new, I encourage you to read this chapter. I set the stage for the rest of the book here by showing you what chemistry is, what chemists do, and why you should be interested in chemistry.

I really enjoy chemistry. It's far more than a simple collection of facts and a body of knowledge. I think it's fascinating to watch chemical changes take place, to figure out unknowns, to use instruments, to extend my senses, and to make predictions and figure out why they were right or wrong. It all starts here — with the basics — so welcome to the interesting world of chemistry.

What Exactly Is Chemistry?

Simply put, this whole branch of science is all about *matter,* which is anything that has mass and occupies space. *Chemistry* is the study of the composition and properties of matter and the changes it undergoes.

A lot of chemistry comes into play with that last part — the changes matter undergoes. Matter is made up of either pure substances or mixtures of pure substances. The change from one substance into another is what chemists call a *chemical change,* or *chemical reaction,* and it's a big deal because when it occurs, a brand-new substance is created (see Chapter 2 for the nitty-gritty details).

What is science?

Science is far more than a collection of facts, figures, graphs, and tables. Science is a method for examining the physical universe. It's a way of asking and answering questions. Science is best described by the attitudes of scientists themselves: They're skeptical — they must be able to test phenomena. And they hold onto the results of their experiments tentatively, waiting for another scientist to disprove them. If it can't be tested, it's not science. Scientists wonder, they question, they strive to find out *why*, and they experiment — they have exactly the same attitudes that most small children have before they grow up. Maybe this is a good definition of scientists — they are adults who've never lost that wonder of nature and the desire to know.

Branches in the tree of chemistry

The general field of chemistry is so huge that it was originally subdivided into a number of different areas of specialization. But there's now a tremendous amount of overlap between the different areas of chemistry, just as there is among the various sciences. Here are the traditional fields of chemistry:

- **Analytical chemistry:** This branch is highly involved in the analysis of substances. Chemists from this field of chemistry may be trying to find out what substances are in a mixture *(qualitative analysis)* or how much of a particular substance is present *(quantitative analysis)* in something. A lot of instrumentation is used in analytical chemistry.
- **Biochemistry:** This branch specializes in living organisms and systems. Biochemists study the chemical reactions that occur at the *molecular level* of an organism — the level where items are so small that people can't directly see them. Biochemists study processes such as digestion, metabolism, reproduction, respiration, and so on. Sometimes it's difficult to distinguish between a biochemist and a molecular biologist because they both study living systems at a microscopic level. However, a biochemist really concentrates more on the reactions that are occurring.
- **Biotechnology:** This is a relatively new area of science that is commonly placed with chemistry. It's the application of biochemistry and biology when creating or modifying genetic material or organisms for specific purposes. It's used in such areas as cloning and the creation of disease-resistant crops, and it has the potential for eliminating genetic diseases in the future.
- **Inorganic chemistry:** This branch is involved in the study of inorganic compounds such as salts. It includes the study of the structure and properties of these compounds. It also commonly involves the study of the individual elements of the compounds. Inorganic chemists would probably say that it is the study of everything except carbon, which they leave to the organic chemists.

So what are compounds and elements? Just more of the anatomy of matter. Matter is made up of either pure substances or mixtures of pure substances, and substances themselves are made up of either elements or compounds. (Chapter 2 dissects the anatomy of matter. And, as with all matters of dissection, it's best to be prepared — with a nose plug and an empty stomach.)

- **Organic chemistry:** This is the study of carbon and its compounds. It's probably the most organized of the areas of chemistry — with good reason. There are millions of organic compounds, with thousands more discovered or created each year. Industries such as the polymer industry, the petrochemical industry, and the pharmaceutical industry depend on organic chemists.
- **Physical chemistry:** This branch figures out how and why a chemical system behaves as it does. Physical chemists study the physical properties and behavior of matter and try to develop models and theories that describe this behavior.

The scientific method

Scientific method is normally described as the way scientists go about examining the physical world around them. In fact, there is no one scientific method that everyone uses every time, but the one I cover here describes most of the critical steps scientists go through sooner or later.

Scientists make observations and note facts regarding something in the physical universe. The observations may raise a question or problem that the researcher wants to solve. He or she comes up with a *hypothesis,* a tentative explanation that's consistent with the observations. The researcher then designs an *experiment* to test the hypothesis. This experiment generates observations or facts that can then be used to generate another hypothesis or modify the current one. Then more experiments are designed, and the loop continues.

In good science, this loop never ends. As scientists become more sophisticated in their scientific skills and build better and better instruments, their hypotheses are tested over and over. But a couple of things can come out of this loop. First, a law may be created. A *law* is a generalization of *what* happens in the scientific system being studied. And like the laws that have been created for the judicial system, scientific laws sometimes have to be modified based on new facts. A theory or model may also be proposed. A *theory* or *model* attempts to explain *why* something happens. It's similar to a hypothesis except that it has much more evidence to support it. The power of the theory or model is prediction. If the scientist can use the model to gain a good understanding of the system, then he or she can make predictions based on the model and then check them out with more experimentation. The observations from this experimentation can be used to refine or modify the theory or model, thus establishing another loop in the process. When does it end? Never.

Macroscopic versus microscopic viewpoints

Most chemists that I know operate quite comfortably in two worlds. One is the *macroscopic* world that you and I see, feel, and touch. This is the world of stained lab coats — of weighing out things like sodium chloride to create things like hydrogen gas. This is the world of experiments, or what some nonscientists call the "real world."

But chemists also operate quite comfortably in the *microscopic* world that you and I can't directly see, feel, or touch. Here, chemists work with theories and models. They may measure the volume and pressure of a gas in the macroscopic world, but they have to mentally translate the measurements into how close the gas particles are in the microscopic world.

Scientists often become so accustomed to slipping back and forth between these two worlds that they do so without even realizing it. An occurrence or observation in the macroscopic world generates an idea related to the microscopic world, and vice versa. You may find this flow of ideas disconcerting at first. But as you study chemistry, you'll soon adjust so that it becomes second nature.

Pure versus applied chemistry

In *pure chemistry,* chemists are free to carry out whatever research interests them — or whatever research they can get funded. There is no real expectation of practical application at this point. The researcher simply wants to know for the sake of knowledge. This type of research (often called *basic research*) is most commonly conducted at colleges and universities. The chemist uses undergraduate and graduate students to help conduct the research. The work becomes part of the professional training of the student. The researcher publishes his or her results in professional journals for other chemists to examine and attempt to refute. Funding is almost always a problem, because the experimentation, chemicals, and equipment are quite expensive.

In *applied chemistry,* chemists normally work for private corporations. Their research is directed toward a very specific short-term goal set by the company — product improvement or the development of a disease-resistant strain of corn, for example. Normally, more money is available for equipment and instrumentation with applied chemistry, but there's also the pressure of meeting the company's goals.

These two types of chemistry, pure and applied, share the same basic differences as science and technology. In *science,* the goal is simply the basic acquisition of knowledge. There doesn't need to be any apparent practical application. Science is simply knowledge for knowledge's sake. *Technology* is the application of science toward a very specific goal.

There's a place in our society for science *and* technology — likewise for the two types of chemistry. The pure chemist generates data and information that is then used by the applied chemist. Both types of chemists have their own sets of strengths, problems, and pressures. In fact, because of the dwindling federal research dollars, many universities are becoming much more involved in gaining patents, and they're being paid for technology transfers into the private sector.

So What Does a Chemist Do All Day?

You can group the activities of chemists into these major categories:

- **Chemists analyze substances.** They determine what is in a substance, how much of something is in a substance, or both. They analyze solids, liquids, and gases. They may try to find the active compound in a substance found in nature, or they may analyze water to see how much lead is present.
- **Chemists create, or *synthesize,* new substances.** They may try to make the synthetic version of a substance found in nature, or they may create an entirely new and unique compound. They may try to find a way to synthesize insulin. They may create a new plastic, pill, or paint. Or they may try to find a new, more efficient process to use for the production of an established product.
- **Chemists create models and test the predictive power of theories.** This area of chemistry is referred to as *theoretical chemistry.* Chemists who work in this branch of chemistry use computers to model chemical systems. Theirs is the world of mathematics and computers. Some of these chemists don't even own a lab coat.
- **Chemists measure the physical properties of substances.** They may take new compounds and measure the melting points and boiling points. They may measure the strength of a new polymer strand or determine the octane rating of a new gasoline.

And Where Do Chemists Actually Work?

You may be thinking that all chemists can be found deep in a musty lab, working for some large chemical company, but chemists hold a variety of jobs in a variety of places:

- **Quality control chemist:** These chemists analyze raw materials, intermediate products, and final products for purity to make sure that they fall within specifications. They may also offer technical support for the customer or analyze returned products. Many of these chemists often solve problems when they occur within the manufacturing process.

- **Industrial research chemist:** Chemists in this profession perform a large number of physical and chemical tests on materials. They may develop new products, and they may work on improving existing products. They may work with particular customers to formulate products that meet specific needs. They may also supply technical support to customers.
- **Sales representative:** Chemists may work as sales representatives for companies that sell chemicals or pharmaceuticals. They may call on their customers and let them know of new products being developed. They may also help their customers solve problems.
- **Forensic chemist:** These chemists may analyze samples taken from crime scenes or analyze samples for the presence of drugs. They may also be called to testify in court as expert witnesses.
- **Environmental chemist:** These chemists may work for water purification plants, the Environmental Protection Agency, the Department of Energy, or similar agencies. This type of work appeals to people who like chemistry but also like to get out in nature. They often go out to sites to collect their own samples.
- **Preservationist of art and historical works:** Chemists may work to restore paintings or statues, or they may work to detect forgeries. With air and water pollution destroying works of art daily, these chemists work to preserve our heritage.
- **Chemical educator:** Chemists working as educators may teach physical science and chemistry in public schools. They may also teach at the college or university level. University chemistry teachers often conduct research and work with graduate students. Chemists may even become chemical education specialists for organizations such as the American Chemical Society.

These are just a few of the professions chemists may find themselves in. I didn't even get into law, medicine, technical writing, governmental relations, and consulting. Chemists are involved in almost every aspect of society. Some chemists even write books.

If you aren't interested in becoming a chemist, why should you be interested in chemistry? (The quick answer is probably "to pass a course.") Chemistry is an integral part of our everyday world, and knowing something about chemistry will help you interact more effectively with our chemical environment.

Chapter 2

Matter and Energy

In This Chapter

- Understanding the states of matter and their changes
- Differentiating between pure substances and mixtures
- Finding out about the metric system
- Examining the properties of chemical substances
- Discovering the different types of energy
- Measuring the energy in chemical bonds

Walk into a room and turn on the light. Look around — what do you see? There might be a table, some chairs, a lamp, a computer humming away. But really all you see is matter and energy. There are many kinds of matter and many kinds of energy, but when all is said and done, you're left with these two things. Scientists used to believe that these two were separate and distinct, but now they realize that matter and energy are linked. In an atomic bomb or nuclear reactor, matter is converted into energy. Perhaps someday the science fiction of *Star Trek* will become a reality and converting the human body into energy and back in a transporter will be commonplace. But in the meantime, I'll stick to the basics of matter and energy.

In this chapter, I cover the two basic components of the universe — matter and energy. I examine the different states of matter and what happens when matter goes from one state to another. I show you how the metric system is used to make matter and energy measurements, and I examine the different types of energy and see how energy is measured.

States of Matter: Macroscopic and Microscopic Views

Look around you. All the stuff you see — your chair, the water you're drinking, the paper this book is printed on — is matter. Matter is the material part of the universe. It's anything that has mass and occupies space. (Later in this chapter, I introduce you to energy, the other part of the universe.) Matter can exist in one of three states: solid, liquid, and gas.

Solids

At the *macroscopic level*, the level at which we directly observe with our senses, a solid has a definite shape and occupies a definite volume. Think of an ice cube in a glass — it's a solid. You can easily weigh the ice cube and measure its volume. At the *microscopic level* (where items are so small that people can't directly observe them), the particles that make up the ice are very close together and aren't moving around very much (see Figure 2-1a).

The reason the particles that make up the ice (also known as *water molecules)* are close together and have little movement is because, in many solids, the particles are pulled into a rigid, organized structure of repeating patterns called a *crystal lattice.* The particles that are contained in the crystal lattice are still moving, but barely — it's more of a slight vibration. Depending on the particles, this crystal lattice may be of different shapes.

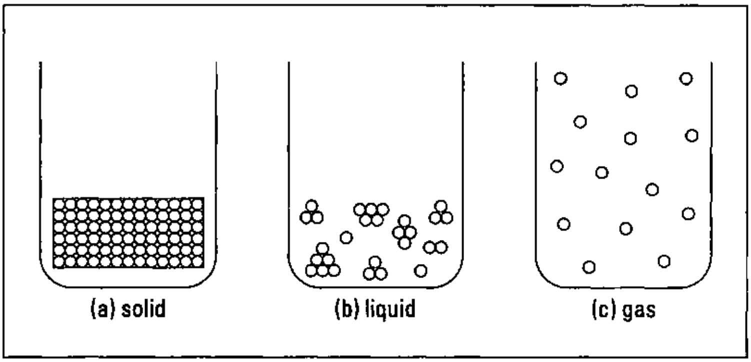


Figure 2-1: Solid, liquid, and gaseous states of matter.

Liquids

When an ice cube melts, it becomes a liquid. Unlike solids, liquids have no definite shape, but they do have a definite volume, just like solids do. For example, a cup of water in a tall skinny glass has a different shape than a cup of water in a pie pan, but in both cases, the volume of water is the same — one cup. Why? The particles in liquids are much farther apart than the particles in solids, and they're also moving around much more (see Figure 2-1b.). Even though the particles are farther apart in liquids than in solids, some particles in liquids may still be near each other, clumped together in small groups.

Because the particles are farther apart in liquids, the attractive forces among them aren't as strong as they are in solids — which is why liquids don't have a definite shape. However, these attractive forces are strong enough to keep the substance confined in one large mass — a liquid — instead of going all over the place.

Gases

If you heat water, you can convert it to *steam,* the gaseous form of water. A gas has no definite shape and no definite volume. In a gas, particles are much farther apart than they are in solids or liquids (see Figure 2-1c), and they're moving relatively independent of each other. Because of the distance between the particles and the independent motion of each of them, the gas expands to fill the area that contains it (and thus it has no definite shape).

Ice in Alaska, Water in Texas: Matter Changes States

When a substance goes from one state of matter to another, we call the process a *change of state.* Some rather interesting things occur during this process.

I'm melting away! Oh, what a world!

Imagine taking a big chunk of ice out of your freezer and putting it into a large pot on your stove. If you measure the temperature of that chunk of ice, you may find it to be –5° Celsius or so. If you take temperature readings while heating the ice, you find that the temperature of the ice begins to rise as the heat from the stove causes the ice particles to begin vibrating faster and faster in the crystal lattice. After a while, some of the particles move so fast that they break free of the lattice, and the crystal lattice (which keeps a solid solid) eventually breaks apart. The solid begins to go from a solid state to a liquid state — a process called *melting.* The temperature at which melting occurs is called the *melting point (mp)* of the substance. The melting point for ice is 32° Fahrenheit, or 0° Celsius.

If you watch the temperature of ice as it melts, you see that the temperature remains steady at 0°C until all the ice has melted. During changes of state (*phase changes*), the temperature remains constant even though the liquid contains more energy than the ice (because the particles in liquids move faster than the particles in solids, as mentioned in the previous section).

Boiling point

If you heat a pot of cool water (or if you continue to heat the pot of now-melted ice cubes mentioned in the preceding section), the temperature of the water rises and the particles move faster and faster as they absorb the heat. The temperature rises until the water reaches the next change of state — boiling. As the particles move faster and faster as they heat up, they begin to break the attractive forces between each other and move freely as steam — a gas. The process by which a substance moves from the liquid state to the gaseous state is called *boiling.* The temperature at which a liquid begins to boil is called the *boiling point (bp).* The bp is dependent on atmospheric pressure, but for water at sea level, it's 212°F, or 100°C. The temperature of the boiling water will remain constant until all the water has been converted to steam.

You can have both water and steam at 100°C. They will have the same temperature, but the steam will have a lot more energy (because the particles move independently and pretty quickly). Because steam has more energy, steam burns are normally a lot more serious than boiling water burns — much more energy is transferred to your skin. I was reminded of this one morning while trying to iron a wrinkle out of a shirt that I was still wearing. My skin and I can attest — steam contains a *lot* of energy!

I can summarize the process of water changing from a solid to a liquid in this way:

ice→water→steam

Because the basic particle in ice, water, and steam is the water molecule (written as H_2O), the same process can also be shown as

$H_2O(s) \rightarrow H_2O(l) \rightarrow H_2O(g)$

Here the *(s)* stands for solid, the *(l)* stands for liquid, and the *(g)* stands for gas. This second depiction is much better, because unlike H_2O, most chemical substances don't have different names for the solid, liquid, and gas forms.

Freezing point: The miracle of ice cubes

If you cool a gaseous substance, you can watch the phase changes that occur. The phase changes are

- **Condensation** — going from a gas to a liquid
- **Freezing** — going from a liquid to a solid

The gas particles have a high amount of energy, but as they're cooled, that energy is reduced. The attractive forces now have a chance to draw the particles closer together, forming a liquid. This process is called *condensation.* The particles are now in clumps (as is characteristic of particles in a liquid state), but as more energy is removed by cooling, the particles start to align themselves, and a solid is formed. This is known as *freezing.* The temperature at which this occurs is called the *freezing point (fp)* of the substance.

The freezing point is the same as the melting point — it's the point at which the liquid is able to become a gas or solid.

I can represent water changing states from a gas to a solid like this:

$$H_2O(g) \rightarrow H_2O(l) \rightarrow H_2O(s)$$

Sublimate this!

Most substances go through the logical progression from solid to liquid to gas as they're heated — or vice versa as they're cooled. But a few substances go directly from the solid to the gaseous state without ever becoming a liquid. Scientists call this process *sublimation.* Dry ice — solid carbon dioxide, written as $CO_2(s)$ — is the classic example of sublimation. You can see dry ice particles becoming smaller as the solid begins to turn into a gas, but no liquid is formed during this phase change. (If you've seen dry ice, then you remember that a white cloud usually accompanies it — magicians and theater productions often use dry ice for a cloudy or foggy effect. The white cloud you normally see isn't the carbon dioxide gas — the gas itself is colorless. The white cloud is the condensation of the water vapor in the air due to the cold of the dry ice.)

The process of sublimation is represented as

$$CO_2(s) \rightarrow CO_2(g)$$

In addition to dry ice, mothballs and certain solid air fresheners also go through the process of sublimation. The reverse of sublimation is *deposition* — going directly from a gaseous state to a solid state.

Pure Substances and Mixtures

One of the basic processes in science is classification. As discussed in the preceding section, chemists can classify matter as solid, liquid, or gas. But there are other ways to classify matter, as well. In this section, I discuss how all matter can be classified as either a pure substance or a mixture (see Figure 2-2).

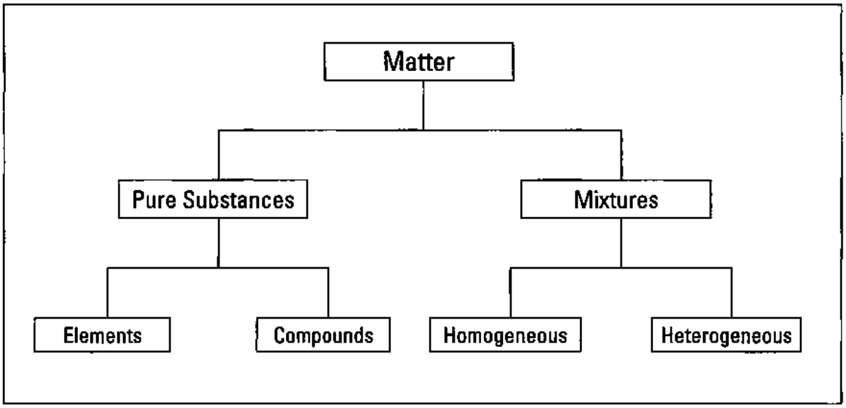


Figure 2-2: Classification of matter.

Pure substances

A *pure substance* has a definite and constant composition or make-up — like salt or sugar.

A pure substance can be either an element or a compound, but the composition of a pure substance doesn't vary.

Elementary, my dear reader

An *element* is composed of a single kind of atom. An *atom* is the smallest particle of an element that still has all the properties of the element. Here's an example: Gold is an element. If you slice and slice a chunk of gold until only one tiny particle is left that can't be chopped any more without losing the properties that make gold *gold,* then you've got an atom.

The atoms in an element all have the same number of protons. *Protons* are subatomic particles — particles of an atom. There are three major subatomic particles, which Chapter 3 covers in great, gory detail.

The important thing to remember right now is that elements are the building blocks of matter. And they're represented in a strange table you may have seen at one time or another — the periodic table. (If you haven't seen such a table before, it's just a list of elements. Chapter 3 contains one if you want to take a peek.)

Compounding the problem

A *compound* is composed of two or more elements in a specific ratio. For example, water (H_2O) is a compound made up of two elements, hydrogen (H) and oxygen (O). These elements are combined in a very specific way — in a ratio of two hydrogen atoms to one oxygen atom (hence H_20). A lot of compounds contain hydrogen and oxygen, but only one has that special 2 to 1 ratio we call water. Even though water is made up of hydrogen and oxygen, the compound water has physical and chemical properties different from both hydrogen and oxygen — water's properties are a unique combination of the two elements.

Chemists can't easily separate the components of a compound: They have to resort to some type of chemical reaction.

Throwing mixtures into the mix

Mixtures are physical combinations of pure substances that have no definite or constant composition — the composition of a mixture varies according to who prepares the mixture. Suppose I asked two people to prepare me a margarita (a delightful mixture). Unless these two people used exactly the same recipe, these mixtures would vary somewhat in their relative amounts of tequila, triple sec, and so on. They would have produced two slightly different mixtures. However, each component of a mixture (that is, each pure substance that makes up the mixture — in the drink example, each *ingredient*) retains its own set of physical and chemical characteristics. Because of this, it's relatively easy to separate the various substances in a mixture.

Although chemists have a difficult time separating compounds into their specific elements, the different parts of a mixture can be easily separated by physical means, such as filtration. For example, suppose you have a mixture of salt and sand, and you want to purify the sand by removing the salt. You can do this by adding water, dissolving the salt, and then filtering the mixture. You then end up with pure sand.

Mixtures can be either homogeneous or heterogeneous.

Homogeneous mixtures, sometimes called *solutions,* are relatively uniform in composition; every portion of the mixture is like every other portion. If you dissolve sugar in water and mix it really well, your mixture is basically the same no matter where you sample it.

But if you put some sugar in a jar, add some sand, and then give the jar a couple of shakes, your mixture doesn't have the same composition throughout the jar. Because the sand is heavier, there's probably more sand at the bottom of the jar and more sugar at the top. In this case, you have a *heterogeneous mixture,* a mixture whose composition varies from position to position within the sample.

Measuring Matter

Scientists are often called on to make measurements, which may include such things as mass (weight), volume, and temperature. If each nation had its own measurement system, communication among scientists would be tremendously hampered, so a worldwide measurement system has been adopted to ensure that scientists can speak the same language.

The SI system

The *SI system* (from the French *Systeme International*) is a worldwide measurement system based on the older metric system that most of us learned in school. There are minor differences between the SI and metric systems, but, for purposes of this book, they're interchangeable.

SI is a decimal system with basic units for things like mass, length, and volume, and prefixes that modify the basic units. For example, the prefix *kilo- (k)* means 1,000. So a kilogram (kg) is 1,000 grams and a kilometer (km) is 1,000 meters. Two other very useful SI prefixes are *centi- (c)* and *milli- (m),* which mean 0.01 and 0.001, respectively. So a milligram (mg) is 0.001 grams — or you can say that there are 1,000 milligrams in a gram. (Check out Appendix A for the most useful SI prefixes.)

SI/English conversions

Many years ago, there was a movement in the United States to convert to the metric system. But, alas, Americans are still buying their potatoes by the pound and their gasoline by the gallon. Don't worry about it. Most professional chemists I know use both the U.S. and SI systems without any trouble. It's necessary to make conversions when using two systems, but I show you how to do that right here.

The basic unit of length in the SI system is the *meter (m).* A meter is a little longer than a yard; there are 1.094 yards in a meter, to be exact. But that's not a really useful conversion. The most useful SI/English conversion for length is

2.54 centimeters = 1 inch

The basic unit of mass in the SI system for chemists is the *gram (g).* And the most useful conversion for mass is

454 grams = 1 pound

The basic unit for volume in the SI system is the *liter (L)*. The most useful conversion is

0.946 liter = 1 quart

By using the preceding conversions and the unit conversion method I describe in Appendix C, you'll be able to handle most SI/English conversions you need to do.

For example, suppose that you have a 5-pound sack of potatoes and you want to know its weight in kilograms. Write down 5 pounds (lbs) as a fraction by placing it over 1.

$$\frac{5.0\,\text{lbs}}{1}$$

Because you need to cancel the unit *lbs* in the numerator, you must find a relationship between *lbs* and something else — and then express that something else with *lbs* in the denominator. You know the relationship between pounds and grams, so you can use that.

$$\frac{5.0\,\cancel{\text{lbs}}}{1} \times \frac{454\,\text{g}}{1\,\cancel{\text{lb}}}$$

Now simply convert from grams to kilograms in the same way.

$$\frac{5.0\,\cancel{\text{lbs}}}{1} \times \frac{454\,\cancel{\text{g}}}{1\,\cancel{\text{lb}}} \times \frac{1\,\text{kg}}{1000\,\cancel{\text{g}}} = 2.3\,\text{kg}$$

Nice Properties You've Got There

When chemists study chemical substances, they examine two types of properties:

- **Chemical properties:** These properties enable a substance to change into a brand-new substance, and they describe how a substance reacts with other substances. Does a substance change into something completely new when water is added — like sodium metal changes to sodium hydroxide? Does it burn in air?
- **Physical properties:** These properties describe the physical characteristics of a substance. The mass, volume, and color of a substance are physical properties, and so is its ability to conduct electricity.

Some physical properties are *extensive properties,* properties that depend on the amount of matter present. Mass and volume are extensive properties. *Intensive properties,* however, don't depend on the amount of matter present. Color is an intensive property. A large chunk of gold, for example, is the same color as a small chunk of gold. The mass and volume of these two chunks are different (extensive properties), but the color is the same. Intensive properties are especially useful to chemists because they can use intensive properties to identify a substance.

How dense are you?

Density is one of the most useful intensive properties of a substance, enabling chemists to more easily identify substances. For example, knowing the differences between the density of quartz and diamond allows a jeweler to check out that engagement ring quickly and easily. *Density (d)* is the ratio of the mass (m) to volume (v) of a substance. Mathematically, it looks like this:

$$d = m/v$$

Usually, mass is described in grams (g) and volume in milliliters (mL), so density is g/mL. Because the volumes of liquids vary somewhat with temperature, chemists also usually specify the temperature at which a density measurement is made. Most reference books report densities at 20°C, because it's close to room temperature and easy to measure without a lot of heating or cooling. The density of water at 20°C, for example, is 1g/mL.

Another term you may sometimes hear is *specific gravity (sg),* which is the ratio of the density of a substance to the density of water at the same temperature. Specific gravity is just another way for you to get around the problem of volumes of liquids varying with the temperature. Specific gravity is used with urinalysis in hospitals and to describe automobile battery fluid in auto repair shops. Note that specific gravity has no units of measure associated with it, because the units g/mL appear in both the numerator and denominator, canceling each other out (see the "SI/English conversions" section, earlier in this chapter, for info about canceling out units of measure). In most cases, the density and specific gravity are almost the same, so it's common to simply use the density.

TIP

You may sometimes see density reported as g/cm^3 or g/cc. These examples are the same as g/mL. A cube measuring 1 centimeter on each edge (written as 1 cm^3) has a volume of 1 milliliter (1 mL). Because 1 mL = 1 cm^3, g/mL and g/cm^3 are interchangeable. And because a cubic centimeter (cm^3) is commonly abbreviated *cc,* g/cc also means the same thing. (You hear *cc* a lot in the medical profession. When you receive a 10cc injection, you're getting 10 milliliters of liquid.)

Measuring density

Calculating density is pretty straightforward. You measure the mass of an object by using a balance or scale, determine the object's volume, and then divide the mass by the volume.

Determining the volume of liquids is easy, but solids can be tricky. If the object is a regular solid, like a cube, you can measure its three dimensions and calculate the volume by multiplying the length by the width by the height (volume = l × w × h). But if the object is an irregular solid, like a rock, determining the volume is more difficult. With irregular solids, you can measure the volume by using something called the Archimedes Principle.

The *Archimedes Principle* states that the volume of a solid is equal to the volume of water it displaces. The Greek mathematician Archimedes discovered this concept in the third century B.C., and finding an object's density is greatly simplified by using it. Say that you want to measure the volume of a small rock in order to determine its density. First, put some water into a graduated cylinder with markings for every mL and read the volume. (The example in Figure 2-3 shows 25 mL.) Next, put the rock in, making sure that it's totally submerged, and read the volume again (29 mL in Figure 2-3). The difference in volume (4 mL) is the volume of the rock.

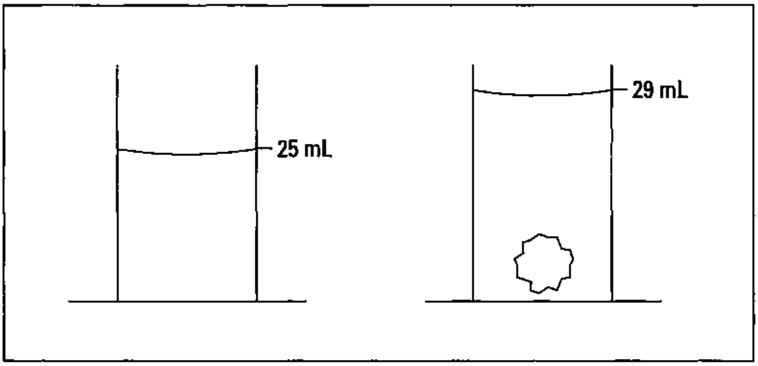


Figure 2-3: Determining the volume of an irregular solid: The Archimedes Principle.

TIP

Anything with a density lower than water will float when put into water, and anything with a density greater than 1 g/mL will sink.

For your pondering pleasure, Table 2-1 lists the density of some common materials.

Table 2-1 Densities of Typical Solids and Liquids in g/mL

Substance	*Density*
Gasoline	0.68
Ice	0.92
Water	1.00
Table Salt	2.16
Iron	7.86
Lead	11.38
Mercury	13.55
Gold	19.3

Energy (Wish I Had More)

Matter is one of two components of the universe. Energy is the other. *Energy* is the ability to do work. And if you're like I am, at about 5 p.m. your ability to do work — and your energy level — are pretty low.

Energy can take several forms — such as heat energy, light energy, electrical energy, and mechanical energy. But two general categories of energy are especially important to chemists — kinetic energy and potential energy.

Kinetic energy — moving right along

Kinetic energy is energy of motion. A baseball flying through the air toward a batter has a large amount of kinetic energy. Ask anyone who's ever been hit with a baseball, and I'm sure that they'll agree! Chemists sometimes study moving particles, especially gases, because the kinetic energy of these particles helps determine whether a particular reaction may take place. The reason is that collisions between particles and the transfer of energy cause chemical reactions to occur.

The kinetic energy of moving particles can be transferred from one particle to another. Have you ever shot pool? You transfer kinetic energy from your moving pool stick to the cue ball to (hopefully) the ball you're aiming at.

Kinetic energy can be converted into other types of energy. In a hydroelectric dam, the kinetic energy of the falling water is converted into electrical energy. In fact, a scientific law — *The Law of Conservation of Energy* — states that in ordinary chemical reactions (or physical processes), energy is neither created nor destroyed but can be converted from one form to another. (This law doesn't hold in nuclear reactions, though. Chapter 5 tells you why.)

Potential energy — sitting pretty

Suppose you take a ball and throw it up into a tree where it gets stuck. You gave that ball kinetic energy — energy in motion — when you threw it. But where's that energy now? It's been converted into the other major category of energy — potential energy.

Potential energy is stored energy. Objects may have potential energy stored in terms of their position. That ball up in the tree has potential energy due to its height. If the ball were to fall, that potential energy would be converted to kinetic energy. (Watch out!)

Potential energy due to position isn't the only type of potential energy. In fact, chemists really aren't all that interested in potential energy due to position. Chemists are far more interested in the energy stored (potential energy) in *chemical bonds,* which are the forces that hold atoms together in compounds.

It takes a lot of energy to run a human body. What if there were no way to store the energy you extract from food? You'd have to eat all the time just to keep your body going. (My wife claims I eat all the time, anyway!) But humans can store energy in terms of chemical bonds. And then later, when we need that energy, our bodies can break those bonds and release it.

The same is true of the fuels we commonly use to heat our homes and run our automobiles. Energy is stored in these fuels — gasoline, for example — and is released when chemical reactions take place.

Measuring Energy

Measuring potential energy can be a difficult task. The potential energy of a ball stuck up in a tree is related to the mass of the ball and its height above the ground. The potential energy contained in chemical bonds is related to the type of bond and the number of bonds that can potentially break.

It's far easier to measure kinetic energy. You can do that with a relatively simple instrument — a thermometer.

Temperature and temperature scales

When you measure, say, the air temperature in your backyard, you're really measuring the average kinetic energy (the energy of motion) of the gas particles in your backyard. The faster those particles are moving, the higher the temperature is.

Now all the particles aren't moving at the same speed. Some are going very fast, and some are going relatively slow, but most are moving at a speed between the two extremes. The temperature reading from your thermometer is related to the *average* kinetic energy of the particles.

You probably use the Fahrenheit scale to measure temperatures, but most scientists and chemists use either the Celsius (°C) or Kelvin (K) temperature scale. (There's no degree symbol associated with K.) Figure 2-4 compares the three temperature scales using the freezing point and boiling point of water as reference points.

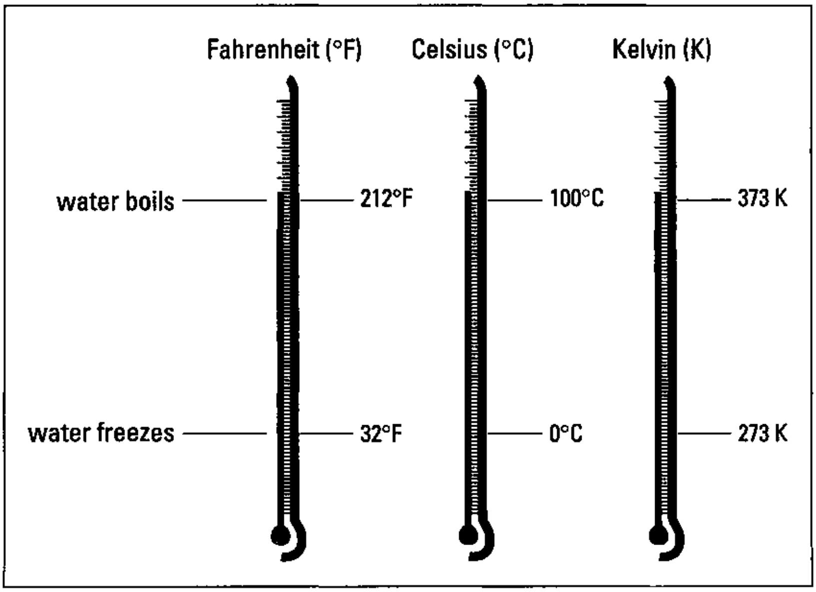


Figure 2-4: Comparison of the Fahrenheit, Celsius, and Kelvin temperature scales.

As you can see from Figure 2-4, water boils at 100°C (373K) and freezes at 0°C (273K). To get the Kelvin temperature, you take the Celsius temperature and add 273. Mathematically, it looks like this:

$$K = {}^\circ C + 273$$

You may want to know how to convert from Fahrenheit to Celsius (because most of us still think in °F). Here are the equations you need:

$${}^\circ C = 5/9\ ({}^\circ F - 32)$$

Be sure to subtract 32 from your Fahrenheit temperature before multiplying by 5/9.

$${}^\circ F = 9/5\ ({}^\circ C) + 32$$

Be sure to multiply your Celsius temperature by 9/5 and *then* add 32.

Go ahead — try these equations out by confirming that the normal body temperature of 98.6°F equals 37°C.

Most of the time in this book, I use the Celsius scale. But when I describe the behavior of gases, I use the Kelvin scale.

Feel the heat

Heat is not the same as temperature. When you measure the temperature of something, you're measuring the average kinetic energy of the individual particles. *Heat,* on the other hand, is a measure of the total amount of energy a substance possesses. For example, a glass of water and a swimming pool may be the same temperature, but they contain vastly different amounts of heat. It takes much more energy to raise the temperature of a swimming pool 5°C than it does a glass of water, because there's so much more water in the swimming pool.

Counting calories

When you hear the word *calories,* you may think about food and counting calories. Food contains energy (heat). The measure of that energy is the nutritional Calorie (which is commonly capitalized), which is really a kilocalorie (kcal). That candy bar you just ate contained 300 nutritional Calories, which is 300 kcal or 300,000 calories. Thinking of it that way may make it a little easier to resist temptation.

The unit of heat in the SI system is the *joule (J)*. Most of us still use the metric unit of heat, the *calorie (cal)*. Here's the relationship between the two:

1 calorie = 4.184 joule

The calorie is a fairly small amount of heat — the amount it takes to raise the temperature of 1 gram of water 1°C. I often use the *kilocalorie (kcal),* which is 1,000 calories, as a convenient unit of heat. If you burn a large kitchen match completely, it produces about 1 kilocalorie (1,000 cal) of heat.

Chapter 3

Something Smaller Than an Atom? Atomic Structure

In This Chapter

- Taking a look at the particles that make up an atom
- Finding out about isotopes and ions
- Coming to understand electron configurations
- Discovering the importance of valence electrons

I remember learning about atoms as a child in school. My teachers called them building blocks and, in fact, we used blocks and Legos to represent atoms. I also remember being told that atoms were so small that nobody would ever see one. Imagine my surprise years later when the first pictures of atoms appeared. They weren't very detailed, but they did make me stop and think how far science had come. I am still amazed when I see pictures of atoms.

In this chapter, I tell you about atoms, the fundamental building blocks of the universe. I cover the three basic particles of an atom — protons, neutrons, and electrons — and show you where they're located. And I use a slew of pages on electrons themselves, because chemical reactions (where a lot of chemistry comes into play) depend on the loss, gain, or sharing of them.

Subatomic Particles: So That's What's in an Atom

The *atom* is the smallest part of matter that represents a particular element. For quite a while, the atom was thought to be the smallest part of matter that could exist. But in the latter part of the nineteenth century and early part of the twentieth, scientists discovered that atoms are composed of certain subatomic particles and that, no matter what the element, the same subatomic particles make up the atom. The number of the various subatomic particles is the only thing that varies.

Scientists now recognize that there are many subatomic particles (this really makes physicists salivate). But in order to be successful in chemistry, you really only need to be concerned with the three major subatomic particles:

- Protons
- Neutrons
- Electrons

Table 3-1 summarizes the characteristics of these three subatomic particles.

Table 3-1 The Three Major Subatomic Particles

Name	*Symbol*	*Charge*	*Mass (g)*	*Mass (amu)*	*Location*
Proton	p^+	+1	1.673×10^{-24}	1	Nucleus
Neutron	n^0	0	1.675×10^{-24}	1	Nucleus
Electron	e^-	–1	9.109×10^{-28}	0.0005	Outside Nucleus

In Table 3-1, the masses of the subatomic particles are listed in two ways: grams and *amu,* which stands for *atomic mass units.* Expressing mass in amu is much easier than using the gram equivalent.

Atomic mass units are based on something called the Carbon 12 scale, a worldwide standard that's been adopted for atomic weights. By international agreement, a carbon atom that contains 6 protons and 6 neutrons has an atomic weight of exactly 12 amu, so 1 amu is $\frac{1}{12}$ of this carbon atom. I know, what do carbon atoms and the number 12 have to do with anything? Just trust me. Because the mass in grams of protons and neutrons are almost exactly the same, both protons and neutrons are said to have a mass of 1 amu. Notice that the mass of an electron is much smaller than that of either a proton or neutron. It takes almost 2,000 electrons to equal the mass of a single proton.

Table 3-1 also shows the electrical charge associated with each subatomic particle. Matter can be electrically charged in one of two ways: positive or negative. The proton carries one unit of positive charge, the electron carries one unit of negative charge, and the neutron has no charge — it's neutral.

Scientists have discovered through observation that objects with like charges, whether positive or negative, repel each other, and objects with unlike charges attract each other.

The atom itself has no charge. It's neutral. (Well, actually, Chapter 6 explains that certain atoms can gain or lose electrons and acquire a charge. Atoms that gain a charge, either positive or negative, are called *ions.*) So how can an atom be neutral if it contains positively charged protons and negatively charged electrons? Ah, good question. The answer is that there are *equal* numbers of protons and electrons — equal numbers of positive and negative charges — so they cancel each other out.

The last column in Table 3-1 lists the location of the three subatomic particles. Protons and neutrons are located in the *nucleus,* a dense central core in the middle of the atom, while the electrons are located outside the nucleus (see "Where Are Those Electrons?" later in this chapter).

The Nucleus: Center Stage

In 1911, Ernest Rutherford discovered that atoms have a nucleus — a center — containing protons. Scientists later discovered that the nucleus also houses the neutron.

The nucleus is very, very small and very, very dense when compared to the rest of the atom. Typically, atoms have diameters that measure around 10^{-10} meters. (That's small!) Nuclei are around 10^{-15} meters in diameter. (That's *really* small!) For example, if the Superdome in New Orleans represented a hydrogen atom, the nucleus would be about the size of a pea.

The protons of an atom are all crammed together inside the nucleus. Now some of you may be thinking, "Okay, each proton carries a positive charge, and like charges repel each other. So if all the protons are repelling each other, why doesn't the nucleus simply fly apart?" It's *The Force,* Luke. Forces in the nucleus counteract this repulsion and hold the nucleus together. (Physicists call these forces *nuclear glue.* But sometimes this "glue" isn't strong enough, and the nucleus does break apart. This process is called *radioactivity.*)

Not only is the nucleus very small, but it also contains most of the mass of the atom. In fact, for all practical purposes, the mass of the atom is the sum of the masses of the protons and neutrons. (I ignore the minute mass of the electrons unless I'm doing very, very precise calculations.)

The sum of the number of protons plus the number of neutrons in an atom is called the *mass number*. And the number of protons in a particular atom is given a special name, the *atomic number*. Chemists commonly use the symbolization shown in Figure 3-1 to represent these things for a particular element.

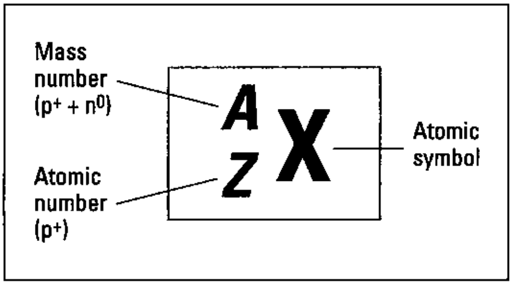


Figure 3-1: Representing a specific element.

As shown in Figure 3-1, chemists use the placeholder *X* to represent the chemical symbol. You can find an element's chemical symbol on the periodic table or a list of elements (see Table 3-2 for a list of elements). The placeholder *Z* represents the atomic number — the number of protons in the nucleus. And *A* represents the mass number, the sum of the number of protons plus neutrons. The mass number (also called the *atomic weight*) is listed in amu.

Suppose you want to represent uranium. You can refer to a periodic table or a list of elements, such as the one shown in Table 3-2, and find that the symbol for uranium is U, its atomic number is 92, and its mass number is 238.

Table 3-2 The Elements

Name	Symbol	Atomic Number	Mass Number	Name	Symbol	Atomic Number	Mass Number
Actinium	Ac	89	227.028	Cerium	Ce	58	140.115
Aluminum	Al	13	26.982	Cesium	Cs	55	132.905
Americium	Am	95	243	Chlorine	Cl	17	35.453
Antimony	Sb	51	121.76	Chromium	Cr	24	51.996
Argon	Ar	18	39.948	Cobalt	Co	27	58.933
Arsenic	As	33	74.922	Copper	Cu	29	63.546
Astatine	At	85	210	Curium	Cm	96	247
Barium	Ba	56	137.327	Dubnium	Db	105	262
Berkelium	Bk	97	247	Dysprosium	Dy	66	162.5
Beryllium	Be	4	9.012	Einsteinium	Es	99	252
Bismuth	Bi	83	208.980	Erbium	Er	68	167.26
Bohrium	Bh	107	262	Europium	Eu	63	151.964
Boron	B	5	10.811	Fermium	Fm	100	257
Bromine	Br	35	79.904	Fluorine	F	9	18.998
Cadmium	Cd	48	112.411	Francium	Fr	87	223
Calcium	Ca	20	40.078	Gadolinium	Gd	64	157.25
Californium	Cf	98	251	Gallium	Ga	31	69.723
Carbon	C	6	12.011	Germanium	Ge	32	72.61

(continued)

Table 3-2 *(continued)*

Name	*Symbol*	*Atomic Number*	*Mass Number*	*Name*	*Symbol*	*Atomic Number*	*Mass Number*
Gold	Au	79	196.967	Mendelevium	Md	101	258
Hafnium	Hf	72	178.49	Mercury	Hg	80	200.59
Hassium	Hs	108	265	Molybdenum	Mo	42	95.94
Helium	He	2	4.003	Neodymium	Nd	60	144.24
Holmfum	Ho	67	164.93	Neon	Ne	10	20.180
Hydrogen	H	1	1.0079	Neptunium	Np	93	237.048
Indium	In	49	114.82	Nickel	Ni	28	58.69
Iodine	I	53	126.905	Niobium	Nb	41	92.906
Iridium	Ir	77	192.22	Nitrogen	N	7	14.007
Iron	Fe	26	55.845	Nobelium	No	102	259
Krypton	Kr	36	83.8	Osmiun	Os	76	190.23
Lanthanum	La	57	138.906	Oxygen	O	8	15.999
Lawrencium	Lr	103	262	Palladium	Pd	46	106.42
Lead	Pb	82	207.2	Phosphorus	P	15	30.974
Lithium	Li	3	6.941	Platinum	Pt	78	195.08
Lutetium	Lu	71	174.967	Plutonium	Pu	94	244
Magnesium	Mg	12	24.305	Polonium	Po	84	209
Manganese	Mn	25	54.938	Potassium	K	19	39.098
Meitnerium	Mt	109	266	Praseodymium	Pr	59	140.908

Name	*Symbol*	*Atomic Number*	*Mass Number*	*Name*	*Symbol*	*Atomic Number*	*Mass Number*
Promethium	Pm	61	145	Tantalum	Ta	73	180.948
Protactinium	Pa	91	231.036	Technetium	Tc	43	98
Radium	Ra	88	226.025	Tellurium	Te	52	127.60
Radon	Rn	86	222	Terbium	Tb	65	158.925
Rhenium	Re	75	186.207	Thallium	Tl	81	204.383
Rhodium	Rh	45	102.906	Thorium	Th	90	232.038
Rubidium	Rb	37	85.468	Thulium	Tm	69	168.934
Ruthenium	Ru	44	101.07	Tin	Sn	50	118.71
Rutherfordium	Rf	104	261	Titanium	Ti	22	47.88
Samarium	Sm	62	150.36	Tungsten	W	74	183.84
Scandium	Sc	21	44.956	Uranium	U	92	238.029
Seaborgium	Sg	106	263	Vanadium	V	23	50.942
Selenium	Se	34	78.96	Xenon	Xe	54	131.29
Silicon	Si	14	28.086	Ytterbium	Yb	70	173.04
Silver	Ag	47	107.868	Yttrium	Y	39	88.906
Sodium	Na	11	22.990	Zinc	Zn	30	65.39
Strontium	Sr	38	87.62	Zirconium	Zr	40	91.224
Sulfur	S	16	32.066				

So you can represent uranium as shown in Figure 3-2.

$$^{238}_{92}\mathrm{U}$$

Figure 3-2: Representing uranium.

You know that uranium has an atomic number of 92 (number of protons) and mass number of 238 (protons plus neutrons). So if you want to know the number of neutrons in uranium, all you have to do is subtract the atomic number (92 protons) from the mass number (238 protons plus neutrons). The resulting number shows that uranium has 146 neutrons.

But how many electrons does uranium have? Because the atom is neutral (it has no electrical charge), there must be equal numbers of positive and negative charges inside it, or equal numbers of protons and electrons. So there are 92 electrons in each uranium atom.

Where Are Those Electrons?

Early models of the atom had electrons spinning around the nucleus in a random fashion. But as scientists learned more about the atom, they found that this representation probably wasn't accurate. Today, two models of atomic structure are used: the Bohr model and the quantum mechanical model. The Bohr model is simple and relatively easy to understand; the quantum mechanical model is based on mathematics and is more difficult to understand. Both, though, are helpful in understanding the atom, so I explain each in the following sections (without resorting to a lot of math).

A model is useful because it helps you understand what's observed in nature. It's not unusual to have more than one model represent and help people understand a particular topic.

The Bohr model — it's really not boring

Have you ever bought color crystals for your fireplace — to make flames of different colors? Or have you ever watched fireworks and wondered where the colors came from?

Color comes from different elements. If you sprinkle table salt — or any salt containing sodium — on a fire, you get a yellow color. Salts that contain copper give a greenish-blue flame. And if you look at the flames through a *spectroscope,* an instrument that uses a prism to break up light into its various components, you see a number of lines of various colors. Those distinct lines of color make up a *line spectrum.*

Niels Bohr, a Danish scientist, explained this line spectrum while developing a model for the atom.

The Bohr model shows that the electrons in atoms are in orbits of differing energy around the nucleus (think of planets orbiting around the sun). Bohr used the term *energy levels* (or *shells*) to describe these orbits of differing energy. And he said that the energy of an electron is *quantized,* meaning electrons can have one energy level or another but nothing in between.

The energy level an electron normally occupies is called its *ground state.* But it can move to a higher-energy, less-stable level, or shell, by absorbing energy. This higher-energy, less-stable state is called the electron's *excited state.*

After it's done being excited, the electron can return to its original ground state by releasing the energy it has absorbed (see Figure 3-3). And here's where the line spectrum explanation comes in. Sometimes the energy released by electrons occupies the portion of the *electromagnetic spectrum* (the range of wavelengths of energy) that humans detect as visible light. Slight variations in the amount of the energy is seen as light of different colors.

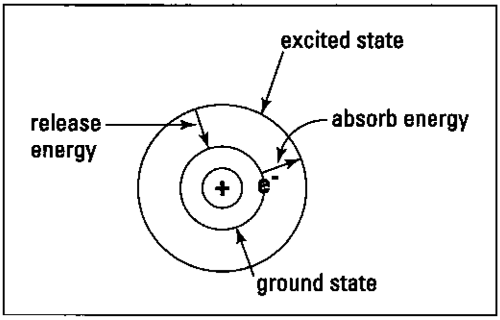


Figure 3-3: Ground and excited states in the Bohr model.

Bohr found that the closer an electron is to the nucleus, the less energy it needs, but the farther away it is, the more energy it needs. So Bohr numbered the electron's energy levels. The higher the energy-level number, the farther away the electron is from the nucleus — and the higher the energy.

Bohr also found that the various energy levels can hold differing numbers of electrons: energy level 1 may hold up to 2 electrons, energy level 2 may hold up to 8 electrons, and so on.

The Bohr model worked well for very simple atoms such as hydrogen (which has 1 electron) but not for more complex atoms. Although the Bohr model is still used today, especially in elementary textbooks, a more sophisticated (and complex) model — the quantum mechanical model — is used much more frequently.

Quantum mechanical model

The simple Bohr model was unable to explain observations made on complex atoms, so a more complex, highly mathematical model of atomic structure was developed — the quantum mechanical model.

This model is based on *quantum theory,* which says matter also has properties associated with waves. According to quantum theory, it's impossible to know the exact position and *momentum* (speed and direction) of an electron at the same time. This is known as the *Uncertainty Principle.* So scientists had to replace Bohr's orbits with *orbitals* (sometimes called *electron clouds*), volumes of space in which there is *likely* to be an electron. In other words, certainty was replaced with probability.

The quantum mechanical model of the atom uses complex shapes of orbitals rather than Bohr's simple circular orbits. Without resorting to a lot of math (you're welcome), this section shows you some aspects of this newest model of the atom.

Four numbers, called *quantum numbers,* were introduced to describe the characteristics of electrons and their orbitals. You'll notice that they were named by totally top-rate techno-geeks:

- Principal quantum number n
- Angular momentum quantum number l
- Magnetic quantum number m_l
- Spin quantum number m_s

Table 3-3 summarizes the four quantum numbers. When they're all put together, theoretical chemists have a pretty good description of the characteristics of a particular electron.

Table 3-3 Summary of the Quantum Numbers

Name	*Symbol*	*Description*	*Allowed Values*
Principal	n	Orbital energy	Positive integers (1, 2, 3, and so on)
Angular momentum	l	Orbital shape	Integers from 0 to $n-1$

Name	*Symbol*	*Description*	*Allowed Values*
Magnetic	m_l	Orientation	Integers from –*l* to 0 to +*l*
Spin	m_s	Electron spin	+½ or –½

The principal quantum number n

The principal quantum number n describes the average distance of the orbital from the nucleus — and the energy of the electron in an atom. It's really about the same as Bohr's energy-level numbers. It can have positive integer (whole number) values: 1, 2, 3, 4, and so on. The larger the value of n, the higher the energy and the larger the orbital. Chemists sometimes call the orbitals *electron shells.*

The angular momentum quantum number l

The angular momentum quantum number *l* describes the shape of the orbital, and the shape is limited by the principal quantum number n: The angular momentum quantum number *l* can have positive integer values from 0 to n–1. For example, if the n value is 3, three values are allowed for *l:* 0, 1, and 2.

The value of *l* defines the shape of the orbital, and the value of n defines the size.

Orbitals that have the same value of n but different values of *l* are called *subshells.* These subshells are given different letters to help chemists distinguish them from each other. Table 3-4 shows the letters corresponding to the different values of *l.*

Table 3-4 Letter Designation of the Subshells

Value of l (subshell)	*Letter*
0	s
1	p
2	d
3	f
4	g

When chemists describe one particular subshell in an atom, they can use both the n value and the subshell letter — 2p, 3d, and so on. Normally, a subshell value of 4 is the largest needed to describe a particular subshell. If chemists ever need a larger value, they can create subshell numbers and letters.

Figure 3-4 shows the shapes of the s, p, and d orbitals.

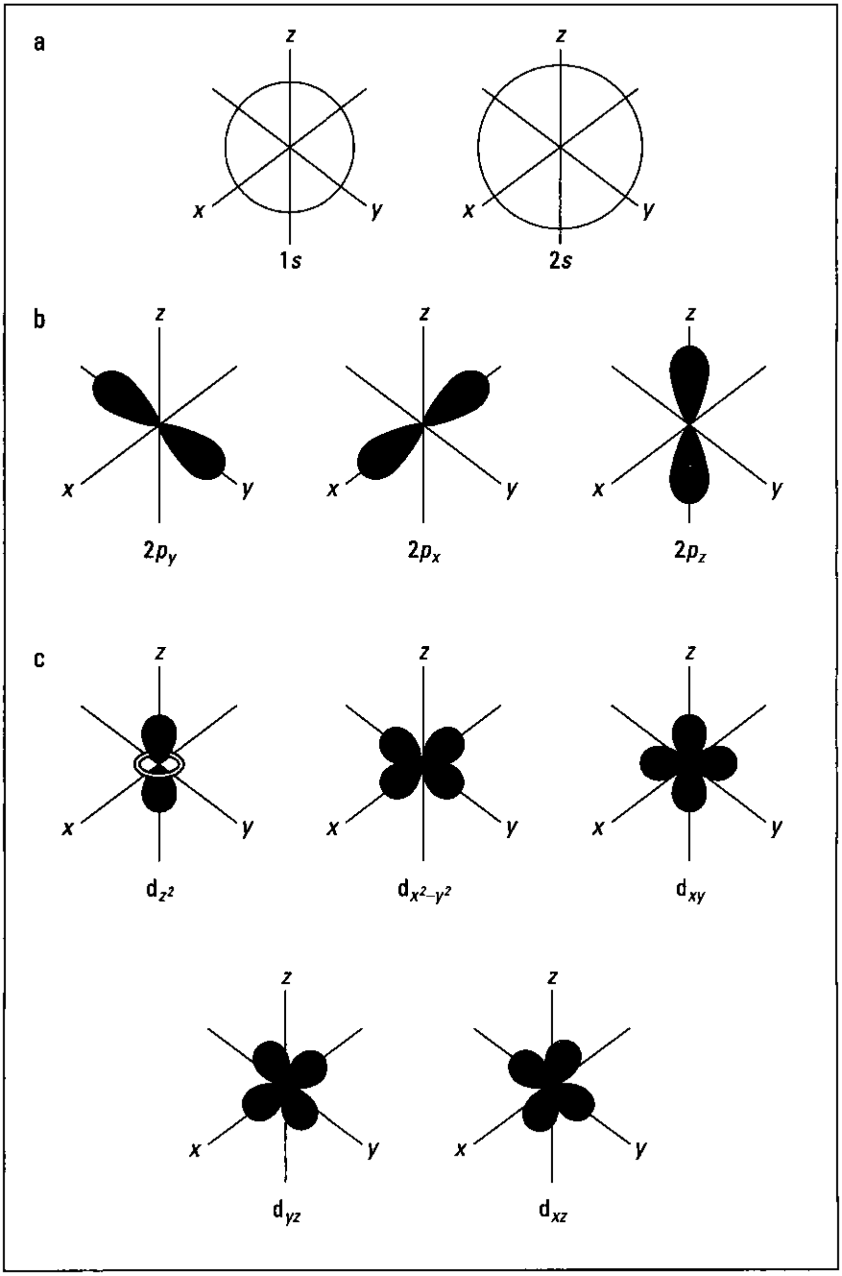


Figure 3-4: Shapes of the s, p, and d orbitals.

In Figure 3-4 (a), there are two s orbitals — one for energy level 1 (1s) and the other for energy level 2 (2s). S orbitals are spherical with the nucleus at the center. Notice that the 2s orbital is larger in diameter than the 1s orbital. In large atoms, the 1s orbital is nestled inside the 2s, just like the 2p is nestled inside the 3p.

Figure 3-4 (b) shows the shapes of the p orbitals, and Figure 3-4 (c) shows the shapes of the d orbitals. Notice that the shapes get progressively more complex.

The magnetic quantum number m_l

The magnetic quantum number m_l describes how the various orbitals are oriented in space. The value of m_l depends on the value of l. The values allowed are integers from $-l$ to 0 to $+l$. For example, if the value of $l = 1$ (p orbital — see Table 3-4), you can write three values for m_l: −1, 0, and +1. This means that there are three different p subshells for a particular orbital. The subshells have the same energy but different orientations in space.

Figure 3-4 (b) shows how the p orbitals are oriented in space. Notice that the three p orbitals correspond to m_l values of −1, 0, and +1, oriented along the x, y, and z axes.

The spin quantum number m_s

The fourth and final (I know you're glad — techie stuff, eh?) quantum number is the spin quantum number m_s. This one describes the direction the electron is spinning in a magnetic field — either clockwise or counterclockwise. Only two values are allowed for m_s: +½ or −½. For each subshell, there can be only two electrons, one with a spin of +½ and another with a spin of −½.

Put all the numbers together and whaddya get? (A pretty table)

I know. Quantum number stuff makes science nerds drool and normal people yawn. But, hey, sometime if the TV's on the blink and you've got some time to kill, take a peek at Table 3-5. You can check out the quantum numbers for each electron in the first two energy levels (oh boy, oh boy, oh boy).

Table 3-5 Quantum Numbers for the First Two Energy Levels

n	l	Subshell Notation	m_l	m_s
1	0	1s	0	+1/2, −1/2
2	0	2s	0	+1/2, −1/2
	1	2p	−1	+1/2, −1/2
			0	+1/2, −1/2
			+1	+1/2, −1/2

Table 3-5 shows that in energy level 1 (n=1) there's only an s orbital. There's no p orbital because an *l* value of 1 (p orbital) is not allowed. And notice that there can be only two electrons in that 1s orbital (m_s of +½ and –½). In fact, there can be only two electrons in any s orbital, whether it's 1s or 5s.

When you move from energy level 1 to energy level 2 (n=2), there can be both s and p orbitals. If you write out the quantum numbers for energy level 3, you see s, p, and d orbitals. Each time you move higher in a major energy level, you add another orbital type.

Notice also that there are three subshells (m*l*) for the 2p orbital (see Figure 3-4 (b)) and that each holds a maximum of two electrons. The three 2p subshells can hold a maximum of six electrons.

There's an energy difference in the major energy levels (energy level 2 is higher in energy than energy level 1), but there's also a difference in the energies of the different orbitals within an energy level. At energy level 2, both s and p orbitals are present. But the 2s is lower in energy than the 2p. The three subshells of the 2p orbital have the same energy. Likewise, the five subshells of the d orbitals (see Figure 3-4 (c)) have the same energy.

Okay. Enough already.

Electron configurations (Bed Check for Electrons)

Chemists find quantum numbers useful when they're looking at chemical reactions and bonding (and those are things many chemists like to study). But they find two other representations for electrons *more* useful and easier to work with:

- ✓ Energy level diagrams
- ✓ Electron configurations

Chemists use both of these things to represent which energy level, subshell, and orbital are occupied by electrons in any particular atom. Chemists use this information to predict what type of bonding will occur with a particular element and show exactly which electrons are being used. These representations are also useful in showing why certain elements behave in similar ways.

In this section, I show you how to use an energy level diagram and write electron configurations.

The dreaded energy level diagram

Figure 3-5 is a blank energy level diagram you can use to depict electrons for any particular atom. Not all the known orbitals and subshells are shown. But with this diagram, you should be able to do most anything you need to. (If you don't have a clue what orbitals, subshells, or all those numbers and letters in the figure have to do with the price of beans, check out the "Quantum mechanical model" section, earlier in this chapter. Fun read, lemme tell ya.)

I represent orbitals with dashes in which you can place a maximum of two electrons. The 1s orbital is closest to the nucleus, and it has the lowest energy. It's also the only orbital in energy level 1 (refer to Table 3-5). At energy level 2, there are both s and p orbitals, with the 2s having lower energy than the 2p. The three 2p subshells are represented by three dashes of the same energy. Energy levels 3, 4, and 5 are also shown. Notice that the 4s has lower energy than the 3d: This is an exception to what you may have thought, but it's what's observed in nature. Go figure. Speaking of which, Figure 3-6 shows the *Aufbau Principle,* a method for remembering the order in which orbitals fill the vacant energy levels.

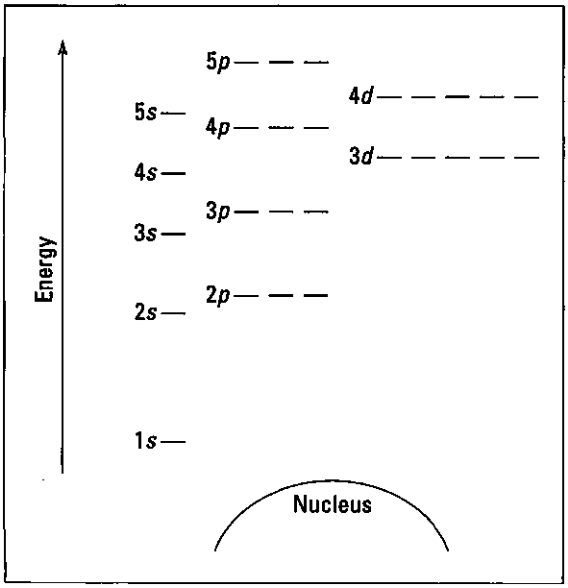


Figure 3-5: Energy level diagram.

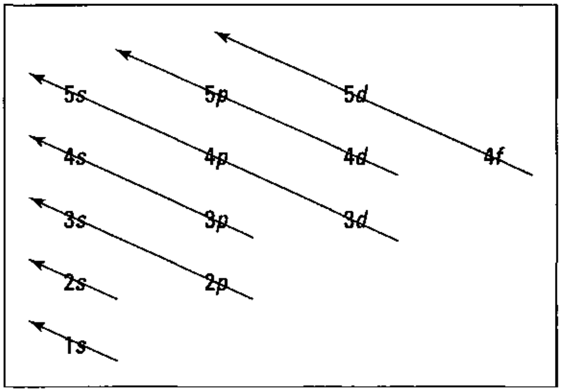


Figure 3-6: Aufbau filling chart.

In using the energy level diagram, remember two things:

- Electrons fill the lowest vacant energy levels first.
- When there's more than one subshell at a particular energy level, such as at the 3p or 4d levels (see Figure 3-5), only one electron fills each subshell until each subshell has one electron. Then electrons start pairing up in each subshell. This rule is named *Hund's Rule.*

Suppose you want to draw the energy level diagram of oxygen. You look on the periodic table or an element list and find that oxygen is atomic number 8. This number means that oxygen has 8 protons in its nucleus and 8 electrons. So you put 8 electrons into your energy level diagram. You can represent electrons as arrows (see Figure 3-7). Note that if two electrons end up in the same orbital, one arrow faces up and the other faces down. (This is called *spin pairing.* It corresponds to the +½ and –½ of m_s.— see "The spin quantum number m_s"section earlier in this chapter.)

The first electron goes into the 1s orbital, filling the lowest energy level first, and the second one spin pairs with the first one. Electrons 3 and 4 spin pair in the next lowest vacant orbital — the 2s. Electron 5 goes into one of the 2p subshells (no, it doesn't matter which one — they all have the same energy), and electrons 6 and 7 go into the other two totally vacant 2p orbitals (see the two things you're supposed to remember about the energy level diagram, a little ways back from here). The last electron spin pairs with one of the electrons in the 2p subshells (again, it doesn't matter which one you pair it with). Figure 3-7 shows the completed energy level diagram for oxygen.

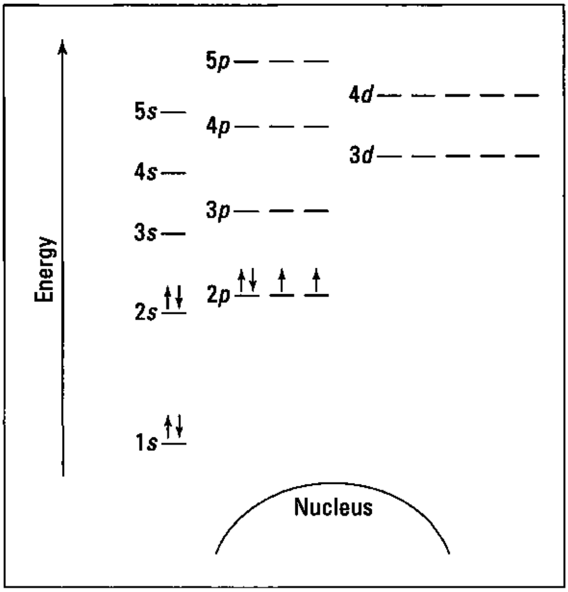


Figure 3-7: Energy level diagram for oxygen.

Electron configurations: Easy and space efficient

Energy level diagrams are useful when you need to figure out chemical reactions and bonding, but they're very bulky to work with. Wouldn't it be nice if there were another representation that gives just about the same information but in a much more concise, shorthand-notation form? Well, there is. It's called the *electron configuration.*

The electron configuration for oxygen is $1s^2 2s^2 2p^4$. Compare that notation with the energy level diagram for oxygen in Figure 3-7. Doesn't the electron configuration take up a lot less space? You can derive the electron configuration from the energy level diagram. The first two electrons in oxygen fill the 1s orbital, so you show it as $1s^2$ in the electron configuration. The 1 is the energy level, the s represents the type of orbital, and the superscript 2 represents the number of electrons in that orbital. The next two electrons are in the 2s orbital, so you write $2s^2$. And, finally, you show the 4 electrons in the 2p orbital as $2p^4$. Put it all together, and you get $1s^2 2s^2 2p^4$.

Some people use a more expanded form, showing how the individual p_x, p_y, and p_z orbitals are oriented along the x,y, and z axes and the number of electrons in each orbital. (The section "The magnetic quantum number m_l," earlier in this chapter, explains how orbitals are oriented in space.) The expanded form is nice if you're really looking at the finer details, but most of the time you won't need that amount of detail in order to show bonding situations and such, so I'm not going to explain the expanded form here.

The sum of the superscript numbers equals the atomic number, or the number of electrons in the atom.

Here are a couple of electron configurations you can use to check your conversions from energy level diagrams:

Chlorine (Cl): $1s^2 2s^2 2p^6 3s^2 3p^5$
Iron (Fe): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

Although I've showed you how to use the energy level diagram to write the electron configuration, with a little practice, you can omit doing the energy level diagram altogether and simply write the electron configuration by knowing the number of electrons and the orbital filling pattern. Anything to save a little precious time, right?

Valence electrons: Living on the edge

When chemists study chemical reactions, they study the transfer or sharing of electrons. The electrons more loosely held by the nucleus — the electrons in the energy level farthest away from the nucleus — are the ones that are gained, lost, or shared.

Electrons are negatively charged, while the nucleus has a positive charge due to the protons. The protons attract and hold the electrons, but the farther away the electrons are, the less the attractive force.

The electrons in the outermost energy level are commonly called *valence electrons*. Chemists really only consider the electrons in the s and p orbitals in the energy level that is currently being filled as valence electrons. In the electron configuration for oxygen, $1s^2 2s^2 2p^4$, energy level 1 is filled, and there are 2 electrons in the 2s orbital and 4 electrons in the 2p orbital for a total of 6 valence electrons. Those valence electrons are the ones lost, gained, or shared.

Being able to determine the number of valence electrons in a particular atom gives you a big clue as to how that atom will react. In Chapter 4, which gives an overview of the periodic table, I show you a quick way to determine the number of valence electrons without writing the electron configuration of the atom.

Isotopes and Ions: These Are a Few of My Favorite Things

But then again, I'm a nerd. The atoms in a particular element have an identical number of protons and electrons but can have varying numbers of neutrons. If they have different numbers of neutrons, then the atoms are called *isotopes.*

Isolating the isotope

Hydrogen is a common element here on earth. Hydrogen's atomic number is 1 — its nucleus contains 1 proton. The hydrogen atom also has 1 electron. Because it has the same number of protons as electrons, the hydrogen atom is neutral (the positive and negative charges have canceled each other out).

Most of the hydrogen atoms on earth contain no neutrons. You can use the symbolization shown in Figure 3-2 to represent hydrogen atoms that don't contain neutrons, as shown in Figure 3-8 (a).

Figure 3-8: The isotopes of hydrogen.

$^{1}_{1}H$	$^{2}_{1}H$	$^{3}_{1}H$
(a) Hydrogen H – 1	(b) Deuterium H – 2	(c) Tritium H – 3

However, approximately one hydrogen atom out of 6,000 contains a neutron in its nucleus. These atoms are still hydrogen, because they have one proton and one electron; they simply have a neutron that most hydrogen atoms lack. So these atoms are called isotopes. Figure 3-8 (b) shows an isotope of hydrogen, commonly called *deuterium.* It's still hydrogen, because it contains only one proton, but it's different from the hydrogen in Figure 3-8 (a), because it also has one neutron. Because it contains one proton and one neutron, its mass number is two.

There's even an isotope of hydrogen containing two neutrons. This one's called *tritium,* and it's represented in Figure 3-8 (c). Tritium doesn't occur naturally on earth, but it can easily be created.

Now take another look at Figure 3-8. It also shows an alternative way of representing isotopes: Write the element symbol, a dash, and then the mass number.

Now you may be wondering, "If I'm doing a calculation involving the atomic mass of hydrogen, which isotope do I use?" Well, you use an average of all the naturally occurring isotopes of hydrogen. But not a simple average. (You have to take into consideration that there's a *lot* more H-1 than H-2, and you don't even consider H-3, because it's not naturally occurring.) You use a *weighted average,* which takes into consideration the abundances of the naturally occurring isotopes. That's why the atomic mass of hydrogen in Table 3-2 isn't a whole number: It's 1.0079 amu. The number shows that there's a lot more H-1 than H-2.

Many elements have several isotopic forms. You can find out more about them in Chapter 5.

Keeping an eye on ions

Because an atom itself is neutral, throughout this book I say that the number of protons and electrons in atoms are equal. But there are cases in which an atom can acquire an electrical charge. For example, in the compound sodium chloride — table salt — the sodium atom has a positive charge and the chlorine atom has a negative charge. Atoms (or groups of atoms) in which there are unequal numbers of protons and electrons are called *ions.*

The neutral sodium atom has 11 protons and 11 electrons, which means it has 11 positive charges and 11 negative charges. Overall, the sodium atom is neutral, and it's represented like this: Na. But the sodium *ion* contains one more positive charge than negative charge, so it's represented like this: Na^+ (the $^+$ represents its net positive electrical charge).

This unequal number of negative and positive charges can occur in one of two ways: An atom can gain a proton (a positive charge) or lose an electron (a negative charge). So which process is more likely to occur? Well, a rough guideline says that it's easy to gain or lose electrons but very difficult to gain or lose protons.

So atoms become ions by gaining or losing electrons. And ions that have a positive charge are called *cations.* The progression goes like this: The Na^+ ion is formed from the loss of one electron. Because it lost an electron, it has more protons than electrons, or more *positive* charges than negative charges, which means it's now called the Na^+ cation. Likewise, The Mg^{2+} cation is formed when the neutral magnesium atom loses two electrons.

Now consider the chlorine atom in sodium chloride. The neutral chlorine atom has acquired a negative charge by gaining an electron. Because it has unequal numbers of protons and electrons, it's now an ion, represented like this: Cl^-. And because ions that have a negative charge are called *anions,* it's now called the Cl^- anion. (You can get the full scoop on ions, cations, and anions in Chapter 6, if you're interested. This here's just a teaser.)

Just for kicks, here are some extra tidbits about ions for your reading pleasure:

- You can write electron configurations and energy level diagrams for ions. The neutral sodium atom (11 protons) has an electron configuration of $1s^22s^22p^63s^1$. The sodium cation has lost an electron — the valence electron, which is *farthest* away from the nucleus (the 3s electron, in this case). The electron configuration of Na^+ is $1s^22s^22p^6$.
- The electron configuration of the chloride ion (Cl^-) is $1s^22s^22p^63s^23p^6$. This is the same electron configuration as the neutral Argon atom. If two chemical species have the same electron configuration, they're said to be *isoelectronic.* Figuring out chemistry requires learning a whole new language, eh?
- This section has been discussing *monoatomic* (one atom) ions. But *polyatomic* (many atom) ions do exist. The ammonium ion, NH_4^+, is a polyatomic ion, or, specifically, a polyatomic cation. The nitrate ion, NO_3^-, is also a polyatomic ion, or, specifically, a polyatomic anion.
- Ions are commonly found in a class of compounds called *salts,* or *ionic solids.* Salts, when melted or dissolved in water, yield solutions that conduct electricity. A substance that conducts electricity when melted or dissolved in water is called an *electrolyte.* Table salt — sodium chloride — is a good example. On the other hand, when table sugar (sucrose) is dissolved in water, it becomes a solution that doesn't conduct electricity. So sucrose is a *nonelectrolyte.* Whether a substance is an electrolyte or a nonelectrolyte gives clues to the type of bonding in the compound. If the substance is an electrolyte, the compound is probably *ionically bonded* (see Chapter 6). If it's a nonelectrolyte, it's probably *covalently bonded* (see Chapter 7).

Chapter 4

The Periodic Table (But No Chairs)

In This Chapter

- Understanding periodicity
- Figuring out how elements are organized in the periodic table

In this chapter, I introduce you to the second most important tool a chemist possesses — the periodic table. (The most important? The beaker and Bunsen burner he or she brews coffee with.)

Chemists are a little lazy, as are most scientists. They like to put things together into groups based on similar properties. This process, called *classification,* makes it much easier to study a particular system. Scientists have grouped the elements together in the periodic table so they don't have to learn the properties of individual elements. With the periodic table, they can just learn the properties of the various groups. So in this chapter, I show you how the elements are arranged in the table, and I show you some important groups. I also explain how chemists and other scientists go about using the periodic table.

Repeating Patterns of Periodicity

In nature, as well as in things that mankind invents, you may notice some repeating patterns. The seasons repeat their pattern of fall, winter, spring, and summer. The tides repeat their pattern of rising and falling. Tuesday follows Monday, December follows November, and so on. This pattern of repeating order is called *periodicity.*

In the mid-1800s, Dmitri Mendeleev, a Russian chemist, noticed a repeating pattern of chemical properties in the elements that were known at the time. Mendeleev arranged the elements in order of increasing atomic mass (see Chapter 3 for a description of atomic mass), to form something that fairly closely resembles our modern periodic table. He was even able to predict the properties of some of the then-unknown elements. Later, the elements were rearranged in order of increasing *atomic number,* the number of protons in the nucleus of the atom (again, see Chapter 3). Figure 4-1 shows the modern periodic table.

PERIODIC TABLE OF THE ELEMENTS

	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIIIB	9 VIIIB
1	1 H Hydrogen 1.00797								
2	3 Li Lithium 6.939	4 Be Beryllium 9.0122							
3	11 Na Sodium 22.9898	12 Mg Magnesium 24.312							
4	19 K Potassium 39.102	20 Ca Calcium 40.08	21 Sc Scandium 44.956	22 Ti Titanium 47.90	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.9380	26 Fe Iron 55.847	27 Co Cobalt 58.9332
5	37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.905	40 Zr Zirconium 91.22	41 Nb Niobium 92.906	42 Mo Molybdenum 95.94	43 Tc Technetium (99)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.905
6	55 Cs Cesium 132.905	56 Ba Barium 137.34	57 La Lanthanum 138.91	72 Hf Hafnium 179.49	73 Ta Tantalum 180.948	74 W Tungsten 183.85	75 Re Rhenium 186.2	76 Os Osmium 190.2	77 Ir Iridium 192.2
7	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 (269)	109 Mt Meitnerium (268)

Lanthanide Series	58 Ce Cerium 140.12	59 Pr Praseodymium 140.907	60 Nd Neodymium 144.24	61 Pm Promethium (145)	62 Sm Samarium 150.35	63 Eu Europium 151.96
Actinide Series	90 Th Thorium 232.038	91 Pa Protactinium (231)	92 U Uranium 238.03	93 Np Neptunium (237)	94 Pu Plutonium (242)	95 Am Americium (243)

Figure 4-1: The periodic table.

10 VIIIB	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 0
								He Helium 4.0026
			5 B Boron 10.811	6 C Carbon 12.01115	7 N Nitrogen 14.0067	8 O Oxygen 15.9994	9 F Flourine 18.9984	10 Ne Neon 20.183
			13 Al Aluminum 26.9815	14 Si Silicon 28.086	15 P Phosphorus 30.9738	16 S Sulfur 32.064	17 Cl Chlorine 35.453	18 Ar Argon 39.948
28 Ni Nickel 58.71	29 Cu Copper 63.546	30 Zn Zinc 65.37	31 Ga Gallium 69.72	32 Ge Germanium 72.59	33 As Arsenic 74.9216	34 Se Selenium 78.96	35 Br Bromine 79.904	36 Kr Krypton 83.80
46 Pd Palladium 106.4	47 Ag Silver 107.868	48 Cd Cadmium 112.40	49 In Indium 114.82	50 Sn Tin 118.69	51 Sb Antimony 121.75	52 Te Tellurium 127.60	53 I Iodine 126.9044	54 Xe Xenon 131.30
78 Pt Platinum 195.09	79 Au Gold 196.967	80 Hg Mercury 200.59	81 Tl Thallium 204.37	82 Pb Lead 207.19	83 Bi Bismuth 208.980	84 Po Polonium (210)	85 At Astatine (210)	86 Rn Radon (222)
110 Uun Ununnilium (269)	111 Uuu Unununium (272)	112 Uub Ununbium (277)	113 Uut §	114 Uuq Ununquadium (285)	115 Uup §	116 Uuh Ununhexium (289)	117 Uus §	118 Uuo Ununoctium (293)

64 Gd Gadolinium 157.25	65 Tb Terbium 158.924	66 Dy Dysprosium 162.50	67 Ho Holmium 164.930	68 Er Erbium 167.26	69 Tm Thufium 168.934	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.97
96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (254)	100 Fm Fermium (257)	101 Md Mendelevium (258)	102 No Nobelium (259)	103 Lr Lawrencium (260)

§ Note: Elements 113, 115, and 117 are not known at this time, but are included in the table to show their expected positions.

Chemists can't imagine doing much of anything without having access to the periodic table. Instead of learning the properties of 109+ elements (more are created almost every year), chemists — and chemistry students — can simply learn the properties of families of elements, thus saving a lot of time and effort. They can find the relationships among elements and figure out the formulas of many different compounds by referring to the periodic table. The table readily provides atomic numbers, mass numbers, and information about the number of valence electrons.

I remember reading a science fiction story many years ago about an alien life based on the element silicon. Silicon was the logical choice for this story because it's in the same family as carbon, the element that's the basis for life on earth. So the periodic table is an absolute necessity for chemists, chemistry students, and science fiction novelists. Don't leave home without it!

Understanding How Elements Are Arranged in the Periodic Table

Look at the periodic table in Figure 4-1. The elements are arranged in order of increasing atomic number. The atomic number (number of protons) is located right above the element symbol. Under the element symbol is the *atomic mass,* or *atomic weight* (sum of the protons and neutrons). Atomic mass is a *weighted average* of all naturally occurring isotopes. (And if that's Greek to you, just flip to Chapter 3 for tons of fun with atomic mass and isotopes.) Notice also that two rows of elements — Ce-Lu (commonly called the Lanthanides) and Th-Lr (the Actinides) — have been pulled out of the main body of the periodic table. If they were included in the main body of the periodic table, the table would be much larger.

The periodic table is composed of horizontal rows called *periods.* The periods are numbered 1 through 7 on the left-hand side of the table. The vertical columns are called *groups,* or *families.* Members of these families have similar properties (see the section "Families and periods," later in this chapter). The families may be labeled at the top of the columns in one of two ways. The older method uses Roman numerals and letters. Many chemists (especially old ones like me) prefer and still use this method. The newer method simply uses the numbers 1 through 18. I use the older method in describing the features of the table.

Using the periodic table, you can classify the elements in many ways. Two quite useful ways are

- Metals, nonmetals, and metalloids
- Families and periods

Metals, nonmetals, and metalloids

If you look carefully at Figure 4-1, you can see a stair-stepped line starting at Boron (B), atomic number 5, and going all the way down to Polonium (Po), atomic number 84. Except for Germanium (Ge) and Antimony (Sb), all the elements to the left of that line can be classified as *metals.* Figure 4-2 shows the metals.

These metals have properties that you normally associate with the metals you encounter in everyday life. They are solid (with the exception of mercury, Hg, a liquid), shiny, good conductors of electricity and heat, *ductile* (they can be drawn into thin wires), and *malleable* (they can be easily hammered into very thin sheets). And all these metals tend to lose electrons easily (see Chapter 6). As you can see, the vast majority of the elements on the periodic table are classified as metals.

Except for the elements that border the stair-stepped line (more on those in a second), the elements to the right of the line are classified as *nonmetals* (along with hydrogen). These elements are shown in Figure 4-3.

Nonmetals have properties opposite those of the metals. The nonmetals are brittle, not malleable or ductile, poor conductors of both heat and electricity, and tend to gain electrons in chemical reactions. Some nonmetals are liquids.

The elements that border the stair-stepped line are classified as *metalloids,* and they're shown in Figure 4-4.

The metalloids, or *semimetals,* have properties that are somewhat of a cross between metals and nonmetals. They tend to be economically important because of their unique conductivity properties (they only partially conduct electricity), which make them valuable in the semiconductor and computer chip industry. (Did you think the term *silicon valley* referred to a valley covered in sand? Nope. Silicon, one of the metalloids, is used in making computer chips.)

Figure 4-2: The metals.

	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB			IB	IIB				
	(2)	(3)	(4)	(5)	(6)	(7)	(8)	(9)	(10)	(11)	(12)				
3 Li Lithium 6.939	4 Be Beryllium 9.0122														
11 Na Sodium 22.9898	12 Mg Magnesium 24.312											13 Al Aluminum 26.9815			
19 K Potassium 39.102	20 Ca Calcium 40.08	21 Sc Scandium 44.956	22 Ti Titanium 47.90	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.9380	26 Fe Iron 55.847	27 Co Cobalt 58.9332	28 Ni Nickel 58.71	29 Cu Copper 63.546	30 Zn Zinc 65.37	31 Ga Gallium 69.72			
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.905	40 Zr Zirconium 91.22	41 Nb Niobium 92.906	42 Mo Molybdenum 95.94	43 Tc Technetium (99)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.905	46 Pd Palladium 106.4	47 Ag Silver 107.868	48 Cd Cadmium 112.40	49 In Indium 114.82	50 Sn Tin 118.69		
55 Cs Cesium 132.905	56 Ba Barium 137.34	57 La Lanthanum 138.91	72 Hf Hafnium 179.49	73 Ta Tantalum 180.948	74 W Tungsten 183.85	75 Re Rhenium 186.2	76 Os Osmium 190.2	77 Ir Iridium 192.2	78 Pt Platinum 195.09	79 Au Gold 196.967	80 Hg Mercury 200.59	81 Tl Thallium 204.37	82 Pb Lead 207.19	83 Bi Bismuth 208.980	84 Po Polonium (210)
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 (269)	109 Mt Meitnerium (268)	110 Uun Ununnilium (269)	111 Uuu Unununium (272)	112 Uub Ununbium (277)				

58 Ce Cerium 140.12	59 Pr Praseodymium 140.907	60 Nd Neodymium 144.24	61 Pm Promethium (145)	62 Sm Samarium 150.35	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.924	66 Dy Dysprosium 162.50	67 Ho Holmium 164.930	68 Er Erbium 167.26	69 Tm Thulium 168.934	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.97
90 Th Thorium 232.038	91 Pa Protactinium (231)	92 U Uranium 238.03	93 Np Neptunium (237)	94 Pu Plutonium (242)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (254)	100 Fm Fermium (257)	101 Md Mendelevium (258)	102 No Nobelium (259)	103 Lr Lawrencium (260)

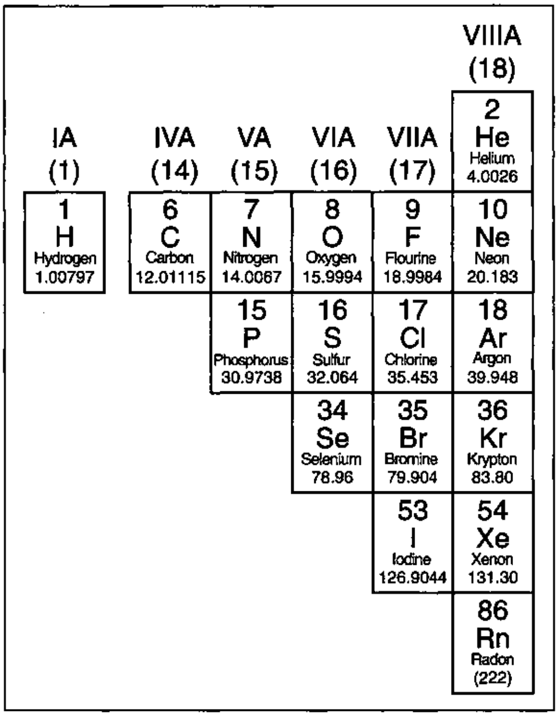


Figure 4-3: The nonmetals.

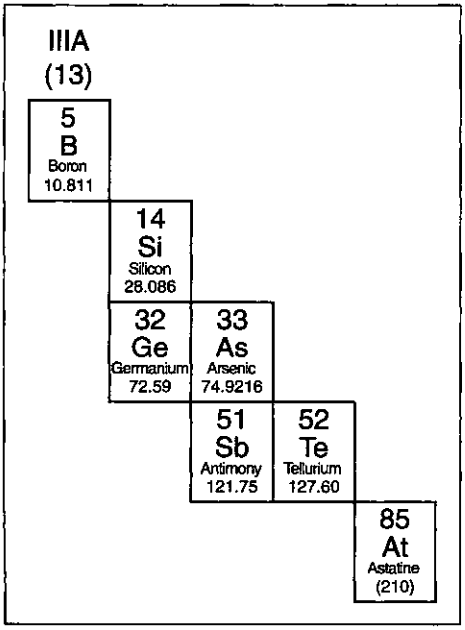


Figure 4-4: The metalloids.

Families and periods

If you refer to the periodic table shown in Figure 4-1, you see seven horizontal rows of elements called periods. In each period, the atomic numbers increase from left to right.

Even though they're in the same period, these elements have chemical properties that are not all that similar. Consider the first two members of period 3: sodium (Na) and magnesium (Mg). In reactions, they both tend to lose electrons (after all, they are metals), but sodium loses one electron, while magnesium loses two. Chlorine (Cl), down near the end of the period, tends to gain an electron (it's a nonmetal). So what you need to remember is that members of a period don't have very similar properties.

The members of a family do have similar properties. Consider the IA family, starting with Lithium (Li) — don't worry about hydrogen, because it's unique, and it doesn't really fit anywhere — and going through Francium (Fr). All these elements tend to lose only one electron in reactions. And all the members of the VIIA family tend to gain one electron.

So why do the elements in the same family have similar properties? And why do some families have the particular properties of electron loss or gain? To find out, you can examine four specific families on the periodic table and look at the electron configurations for a few elements in each family.

My family name is special

Take a look at Figure 4-5, which lists some important families that are given special names:

- The IA family is made up of the *alkali metals.* In reactions, these elements all tend to lose a single electron. This family contains some important elements, such as sodium (Na) and potassium (K). Both of these elements play an important role in the chemistry of the body and are commonly found in salts.
- The IIA family is made up of the *alkaline earth metals.* All these elements tend to lose two electrons. Calcium (Ca) is an important member of the IIA family (you need calcium for healthy teeth and bones).
- The VIIA family is made up of the *halogens.* They all tend to gain a single electron in reactions. Important members in the family include chlorine (Cl), used in making table salt and bleach, and iodine (I). Ever use tincture of iodine as a disinfectant?

- The VIIIA family is made up of the *noble gases.* These elements are *very* unreactive. For a long time, the noble gases were called the inert gases, because people thought that these elements wouldn't react at all. Later, a scientist named Neil Bartlett showed that at least some of the inert gases could be reacted, but they required very special conditions. After Bartlett's discovery, the gases were then referred to as noble gases.

What valence electrons have to do with families

Chapter 3 explains that an *electron configuration* shows the number of electrons in each orbital in a particular atom. The electron configuration forms the basis of the concept of bonding and molecular geometry and other important stuff that I cover in the various chapters of this book.

IA (1)	IIA (2)	VIIA (17)	VIIIA (18)
3 Li Lithium 6.939	4 Be Beryllium 9.0122		2 He Helium 4.0026
11 Na Sodium 22.9898	12 Mg Magnesium 24.312	9 F Flourine 18.9984	10 Ne Neon 20.183
19 K Potassium 39.102	20 Ca Calcium 40.08	17 Cl Chlorine 35.453	18 Ar Argon 39.948
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	35 Br Bromine 79.904	36 Kr Krypton 83.80
55 Cs Cesium 132.905	56 Ba Barium 137.34	53 I Iodine 126.9044	54 Xe Xenon 131.30
87 Fr Francium (223)	88 Ra Radium (226)	85 At Astatine (210)	86 Rn Radon (222)
Alkali Metals	Alkaline Earth Metals	Halogens	Noble Gases

Figure 4-5: Some important chemical families.

Tables 4-1 through 4-4 show the electron configurations for the first three members of the families IA, IIA, VIIA, and VIIIA.

Table 4-1 Electron Configurations for Members of IA (alkali metals)

Element	*Electron Configuration*
Li	$1s^2\mathbf{2s^1}$
Na	$1s^22s^22p^6\mathbf{3s^1}$
K	$1s^22s^22p^63s^23p^6\mathbf{4s^1}$

Table 4-2 Electron Configurations for Members of IIA (alkaline earth metals)

Element	*Electron Configuration*
Be	$1s^2\mathbf{2s^2}$
Mg	$1s^22s^22p^6\mathbf{3s^2}$
Ca	$1s^22s^22p^63s^23p^6\mathbf{4s^2}$

Table 4-3 Electron Configurations for Members of VIIA (halogens)

Element	*Electron Configuration*
F	$1s^2\mathbf{2s^22p^5}$
Cl	$1s^22s^22p^6\mathbf{3s^23p^5}$
Br	$1s^22s^22p^63s^23p^6\mathbf{4s^2}3d^{10}\mathbf{4p^5}$

Table 4-4 Electron Configurations for Members of VIIIA (noble gases)

Element	*Electron Configuration*
Ne	$1s^2\mathbf{2s^22p^6}$
Ar	$1s^22s^22p^6\mathbf{3s^23p^6}$
Kr	$1s^22s^22p^63s^23p^6\mathbf{4s^2}3d^{10}\mathbf{4p^6}$

These electron configurations show that some similarities among each group of elements are in terms of their valence electrons. *Valence electrons* are the s and p electrons in the outermost energy level of an atom (see Chapter 3).

Look at the electron configurations for the alkali metals (Table 4-1). In lithium, energy level 1 is filled, and a single electron is in the 2s orbital. In sodium, energy levels 1 and 2 are filled, and a single electron is in energy level 3. All these elements have one valence electron in an s orbital. The alkaline earth elements (Table 4-2) each have two valence electrons. The halogens (Table 4-3) each have seven valence electrons (in s and p orbitals — d orbitals don't count), and the noble gases (Table 4-4) each have eight valence electrons, which fill their valence orbitals.

So how do you remember all this stuff?

Here's something to keep in mind about the number of valence electrons and the Roman numeral column number: The *I*A family has *1* valence electron; the *II*A family has *2* valence electrons; the *VII*A family has *7* valence electrons; and the *VIII*A family has *8* valence electrons. So for the families labeled with a Roman numeral and an A, the Roman numeral gives the number of valence electrons. Pretty cool, eh?

The Roman numeral makes it very easy to determine that oxygen (O) has six valence electrons (it's in the VIA family), that silicon (Si) has four, and so on. You don't even have to write the electronic configuration or the energy diagram to determine the number of valence electrons.

Noble and gassy

The fact that the noble gases have eight valence electrons, filling their valence, or outermost energy level, explains why the noble gases are extremely hard to react. They are stable, or "satisfied," with a filled (complete) valence energy level. They don't easily lose, gain, or share electrons.

A lot of stability in nature seems to be associated with this condition. Chemists observe that the other elements in the *A* families on the periodic table tend to lose, gain, or share valence electrons in order to achieve the goal of having a filled valence shell of eight electrons: This is sometimes called the *octet rule.* For example, look at the electron configuration for sodium (Na): $1s^22s^22p^63s^1$. It has one valence electron — the $3s^1$. If it lost that electron, its valence shell would be energy level 2, which is filled. Without the $3s^1$, it would become *isoelectronic* (have the same electronic configuration) as Neon (Ne) and achieve stability. As I show you in Chapters 6 and 7, this is the driving force in chemical bonding: achieving stability by having a filled valence shell.

But what about elements that are labeled with a Roman numeral and a B? These elements, found in the middle of the periodic table, are commonly called the *transition metals;* their electrons are progressively filling the d orbitals. Scandium (Sc) is the first member of the transition metals, and it has an electronic configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$. Titanium (Ti), the next transition metal, has a configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$. Notice that the number of electrons in the s and p orbitals is not changing. The progressively added electrons fill the d orbitals. Lanthanides and Actinides, the two groups of elements that are pulled out of the main body of the periodic table and shown below it, are classified as *inner transition metals.* In these elements, the electrons are progressively filling the f orbitals in much the same way that the electrons of the transition metals fill the d orbitals.

Chapter 5

Nuclear Chemistry: It'll Blow Your Mind

In This Chapter

- Understanding radioactivity and radioactive decay
- Figuring out half-lives
- The basics of nuclear fission
- Taking a look at nuclear fusion
- Tracing the effects of radiation

Most of this book deals, in one way or the other, with chemical reactions. And when I talk about these reactions, I'm really talking about how the valence electrons (the electrons in the outermost energy levels of atoms) are lost, gained, or shared. I mention very little about the nucleus of the atom because, to a very large degree, it's not involved in chemical reactions.

But, in this chapter, I do discuss the nucleus and the changes it can undergo. I talk about radioactivity and the different ways an atom can decay. I discuss half-lives and show you why they are important in the storage of nuclear waste products. I also discuss nuclear fission in terms of bombs, power plants, and the hope that nuclear fusion holds for mankind.

Like most of you reading this book, I'm a child of the Atomic Age. I actually remember open air testing of nuclear weapons. I remember being warned not to eat snow because it might contain fallout. I remember friends building fallout shelters. I remember A-bomb drills at school. I remember x-ray machines in shoe stores. (I never did order those x-ray glasses, though!) And I remember radioactive Fiesta stoneware and radium watch hands. When I was growing up, atomic energy was new, exciting, and scary. And it still is.

It All Starts with the Atom

To understand nuclear chemistry, you need to know the basics of atomic structure. Chapter 3 drones on and on (*and on*) about atomic structure, if you're interested. This section just provides a quickie brain dump.

The *nucleus,* that dense central core of the atom, contains both protons and neutrons. Electrons are outside the nucleus in energy levels. Protons have a positive charge, neutrons have no charge, and electrons have a negative charge. A neutral atom contains equal numbers of protons and electrons. But the number of neutrons within an atom of a particular element can vary. Atoms of the same element that have differing numbers of neutrons are called *isotopes.* Figure 5-1 shows the symbolization chemists use to represent a specific isotope of an element.

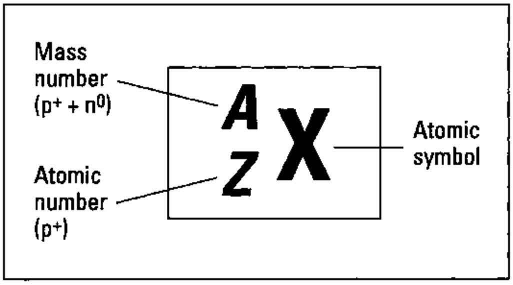


Figure 5-1: Representing a specific isotope.

In the figure, X represents the symbol of the element found on the periodic table, Z represents the *atomic number* (the number of protons in the nucleus), and A represents the *mass number* (the sum of the protons and neutrons in that particular isotope). If you subtract the atomic number from the mass number (A – Z), you get the number of neutrons in that particular isotope. A short way to show the same information is to simply use the element symbol (X) and the mass number (A) — for example, U-235.

Radioactivity and Man-Made Radioactive Decay

For purposes of this book, I define *radioactivity* as the spontaneous decay of an unstable nucleus. An unstable nucleus may break apart into two or more other particles with the release of some energy (see "Gone (Nuclear) Fission," later in this chapter, for more info on this process). This breaking apart can occur in a number of ways, depending on the particular atom that's decaying.

You can often predict one of the particles of a radioactive decay by knowing the other particle. Doing so involves something called *balancing the nuclear reaction.* (A *nuclear reaction* is any reaction involving a change in nuclear structure.)

Balancing a nuclear reaction is really a fairly simple process. But before I explain it, I want to show you how to represent a reaction:

Reactants → Products

Reactants are the substances you start with, and products are the new substances being formed. The arrow, called a *reaction arrow,* indicates that a reaction has taken place.

For a nuclear reaction to be balanced, the sum of all the atomic numbers on the left-hand side of the reaction arrow must equal the sum of all the atomic numbers on the right-hand side of the arrow. The same is true for the sums of the mass numbers. Here's an example: Suppose you're a scientist performing a nuclear reaction by bombarding a particular isotope of chlorine (Cl-35) with a neutron. (Work with me here. I'm just trying to get to a point.) You observe that an isotope of hydrogen, H-1, is created along with another isotope, and you want to figure out what the other isotope is. The equation for this example is

$${}^{35}_{17}Cl+\underset{neutron}{{}^{1}_{0}n}\rightarrow\underline{\text{ Pr? }}+{}^{1}_{1}H$$

Now to figure out the unknown isotope (represented by Pr), you need to balance the equation. The sum of the atomic numbers on the left is 17 (17 + 0), so you want the sum of the atomic numbers on the right to equal 17 too. Right now, you've got an atomic number of 1 on the right; 17 – 1 is 16, so that's the atomic number of the unknown isotope. This atomic number identifies the element as Sulfur (S).

Now look at the mass numbers in the equation. The sum of the mass numbers on the left is 36 (35 + 1), and you want the sum of the mass numbers on the right to equal 36, too. Right now, you've got a mass number of 1 on the right; 36 – 1 is 35, so that's the mass number of the unknown isotope. Now you know that the unknown isotope is a Sulfur isotope (S-35). And here's what the balanced nuclear equation looks like:

$${}^{35}_{17}Cl+\underset{neutron}{{}^{1}_{0}n}\rightarrow{}^{35}_{16}S+{}^{1}_{1}H$$

This equation represents a nuclear *transmutation,* the conversion of one element into another. Nuclear transmutation is a process human beings control. S-35 is an isotope of sulfur that doesn't exist in nature. It's a *man-made isotope.* Alchemists, those ancient predecessors of chemists, dreamed of converting one element into another (usually lead into gold), but they were never able to master the process. Chemists are now able, sometimes, to convert one element into another.

Natural Radioactive Decay: How Nature Does It

Certain isotopes are unstable: Their nucleus breaks apart, undergoing nuclear decay. Sometimes the product of that nuclear decay is unstable itself and undergoes nuclear decay, too. For example, when U-238 (one of the radioactive isotopes of uranium) initially decays, it produces Th-234, which decays to Pa-234. The decay continues until, finally, after a total of 14 steps, Pb-206 is produced. Pb-206 is stable, and the decay sequence, or series, stops.

Before I show you *how* radioactive isotopes decay, I want to briefly explain *why* a particular isotope decays. The nucleus has all those positively charged protons shoved together in an extremely small volume of space. All those protons are repelling each other. The forces that normally hold the nucleus together, the "nuclear glue," sometimes can't do the job, and so the nucleus breaks apart, undergoing nuclear decay.

All elements with 84 or more protons are unstable; they eventually undergo decay. Other isotopes with fewer protons in their nucleus are also radioactive. The radioactivity corresponds to the neutron/proton ratio in the atom. If the neutron/proton ratio is too high (there are too many neutrons or too few protons), the isotope is said to be *neutron rich* and is, therefore, unstable. Likewise, if the neutron/proton ratio is too low (there are too few neutrons or too many protons), the isotope is unstable. The neutron/proton ratio for a certain element must fall within a certain range for the element to be stable. That's why some isotopes of an element are stable and others are radioactive.

There are three primary ways that naturally occurring radioactive isotopes decay:

- Alpha particle emission
- Beta particle emission
- Gamma radiation emission

In addition, there are a couple of less common types of radioactive decay:

- Positron emission
- Electron capture

Alpha emission

An *alpha particle* is defined as a positively charged particle of a helium nuclei. I hear ya: *Huh?* Try this: An alpha particle is composed of two protons and two neutrons, so it can be represented as a Helium-4 atom. As an alpha particle breaks away from the nucleus of a radioactive atom, it has no electrons, so it has a +2 charge. Therefore and to-wit, it's a positively charged particle of a helium nuclei. (Well, it's really a *cation,* a positively charged ion — see Chapter 3.)

But electrons are basically free — easy to lose and easy to gain. So normally, an alpha particle is shown with no charge because it very rapidly picks up two electrons and becomes a neutral helium atom instead of an ion.

Large, heavy elements, such as uranium and thorium, tend to undergo alpha emission. This decay mode relieves the nucleus of two units of positive charge (two protons) and four units of mass (two protons + two neutrons). What a process. Each time an alpha particle is emitted, four units of mass are lost. I wish I could find a diet that would allow me to lose four pounds at a time!

Radon-222 (Rn-222) is another alpha particle emitter, as shown in the following equation:

$$^{222}_{86}Rn \rightarrow {}^{218}_{84}Po + \underbrace{{}^{4}_{2}He}_{\text{alpha particle}}$$

Here, Radon-222 undergoes nuclear decay with the release of an alpha particle. The other remaining isotope must have a mass number of 218 (222 – 4) and an atomic number of 84 (86 – 2), which identifies the element as Polonium (Po). (If this subtraction stuff confuses you, check out how to balance equations in the section "Radioactivity and Man-Made Radioactive Decay," earlier in this chapter.)

Beta emission

A *beta particle* is essentially an electron that's emitted from the nucleus. (Now I know what you're thinking — electrons aren't in the nucleus. Keep on reading to find out how they can be formed in this nuclear reaction.) Iodine-131 (I-131), which is used in the detection and treatment of thyroid cancer, is a beta particle emitter:

$$^{131}_{53}I \rightarrow {}^{131}_{54}Xe + \underbrace{{}^{0}_{-1}e}_{\text{beta particle}}$$

Here, the Iodine-131 gives off a beta particle (an electron), leaving an isotope with a mass number of 131 (131 – 0) and an atomic number of 54 (53 – (-1)). An atomic number of 54 identifies the element as Xenon (Xe).

Notice that the mass number doesn't change in going from I-131 to Xe-131, but the atomic number increases by one. In the iodine nucleus, a neutron was converted (decayed) into a proton and an electron, and the electron was emitted from the nucleus as a beta particle. Isotopes with a high neutron/proton ratio often undergo beta emission, because this decay mode allows the number of neutrons to be decreased by one and the number of protons to be increased by one, thus lowering the neutron/proton ratio.

Gamma emission

Alpha and beta particles have the characteristics of matter: They have definite masses, occupy space, and so on. However, because there is no mass change associated with gamma emission, I refer to gamma emission as *gamma radiation emission.* Gamma radiation is similar to x-rays — high energy, short wavelength radiation. Gamma radiation commonly accompanies both alpha and beta emission, but it's usually not shown in a balanced nuclear reaction. Some isotopes, such as Cobalt-60 (Co-60), give off large amounts of gamma radiation. Co-60 is used in the radiation treatment of cancer. The medical personnel focus gamma rays on the tumor, thus destroying it.

Positron emission

Although positron emission doesn't occur with naturally occurring radioactive isotopes, it does occur naturally in a few man-made ones. A *positron* is essentially an electron that has a positive charge instead of a negative charge. A positron is formed when a proton in the nucleus decays into a neutron and a positively charged electron. The positron is then emitted from the nucleus. This process occurs in a few isotopes, such as Potassium-40 (K-40), as shown in the following equation:

$$^{40}_{19}K \rightarrow {}^{40}_{18}Ar + \underset{\text{positron}}{{}^{0}_{+1}e}$$

The K-40 emits the positron, leaving an element with a mass number of 40 (40 – 0) and an atomic number of 18 (19 – 1). An isotope of argon (Ar), Ar-40, has been formed.

If you watch *Star Trek*, you may have heard about antimatter. The positron is a tiny bit of antimatter. When it comes in contact with an electron, both particles are destroyed with the release of energy. Luckily, not many positrons are produced: If a lot of them were produced, you'd probably have to spend a lot of time ducking explosions.

Electron capture

Electron capture is a rare type of nuclear decay in which an electron from the innermost energy level (the 1s — see Chapter 3) is captured by the nucleus. This electron combines with a proton to form a neutron. The atomic number decreases by one, but the mass number stays the same. The following equation shows the electron capture of Polonium-204 (Po-204):

$$^{204}_{84}Po + ^{0}_{-1}e \rightarrow ^{204}_{83}Bi + x\text{-}rays$$

The electron combines with a proton in the polonium nucleus, creating an isotope of bismuth (Bi-204).

The capture of the 1s electron leaves a vacancy in the 1s orbitals. Electrons drop down to fill the vacancy, releasing energy not in the visible part of the electromagnetic spectrum but in the X-ray portion.

Half-Lives and Radioactive Dating

If you could watch a single atom of a radioactive isotope, U-238, for example, you wouldn't be able to predict when that particular atom might decay. It might take a millisecond, or it might take a century. There's simply no way to tell.

But if you have a large enough sample — what mathematicians call a *statistically significant sample size* — a pattern begins to emerge. It takes a certain amount of time for half the atoms in a sample to decay. It then takes the same amount of time for half the remaining radioactive atoms to decay, and the same amount of time for half of those remaining radioactive atoms to decay, and so on. The amount of time it takes for one-half of a sample to decay is called the *half-life* of the isotope, and it's given the symbol $t_{1/2}$. This process is shown in Table 5-1.

Table 5-1 Half-Life Decay of a Radioactive Isotope

Half-Life	*Percent of Radioactive Isotope Remaining*
0	100.00
1	50.00
2	25.00
3	12.50
4	6.25
5	3.12
6	1.56
7	0.78
8	0.39
9	0.19
10	0.09

It's important to realize that the half-life decay of radioactive isotopes is not linear. For example, you can't find the remaining amount of an isotope as 7.5 half-lives by finding the midpoint between 7 and 8 half-lives. This decay is an example of an exponential decay, shown in Figure 5-2.

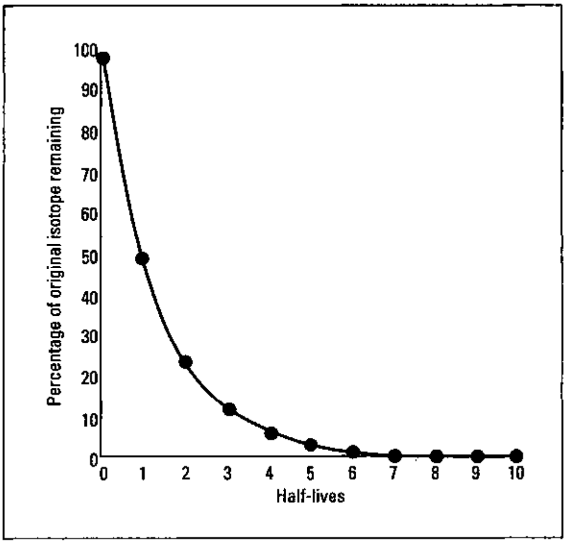


Figure 5-2: Decay of a radioactive isotope.

If you want to find times or amounts that are not associated with a simple multiple of a half-life, you can use this equation:

$$\ln\left(\frac{N_0}{N}\right)=\left(\frac{0.6963}{t_{1/2}}\right)t$$

In the equation, *ln* stands for the *natural logarithm* (not the base 10 log; it's that ln button on your calculator, not the log button), N_o is the amount of radioactive isotope that you start with, N is the amount of radioisotope left at some time *(t)*, and $t_{1/2}$ is the half-life of the radioisotope. If you know the half-life and the amount of the radioactive isotope that you start with, you can use this equation to calculate the amount remaining radioactive at any time. But we are going to keep it simple.

Half-lives may be very short or very long. Table 5-2 shows the half-lives of some typical radioactive isotopes.

Table 5-2 Half-Lives of Some Radioactive Isotopes

Radioisotope	*Radiation Emitted*	*Half-Life*
Kr-94	Beta	1.4 seconds
Rn-222	Alpha	3.8 days
I-131	Beta	8 days
Co-60	Gamma	5.2 years
H-3	Beta	12.3 years
C-14	Beta	5,730 years
U-235	Alpha	4.5 billion years
Re-187	Beta	70 billion years

Safe handling

Knowing about half-lives is important because it enables you to determine when a sample of radioactive material is safe to handle. The rule is that a sample is safe when its radioactivity has dropped below detection limits. And that occurs at 10 half-lives. So, if radioactive iodine-131 ($t_{1/2}$ = 8 days) is injected into the body to treat thyroid cancer, it'll be "gone" in 10 half-lives, or 80 days.

This stuff is important to know when using radioactive isotopes as *medical tracers,* which are taken into the body to allow doctors to trace a pathway or find a blockage, or in cancer treatments. They need to be active long enough to treat the condition, but they should also have a short enough half-life so that they don't injure healthy cells and organs.

Radioactive dating

A useful application of half-lives is *radioactive dating.* No, radioactive dating has nothing to do with taking an X-ray tech to the movies. It has to do with figuring out the age of ancient things.

Carbon-14 (C-14), a radioactive isotope of carbon, is produced in the upper atmosphere by cosmic radiation. The primary carbon-containing compound in the atmosphere is carbon dioxide, and a very small amount of carbon dioxide contains C-14. Plants absorb C-14 during photosynthesis, so C-14 is incorporated into the cellular structure of plants. Plants are then eaten by animals, making C-14 a part of the cellular structure of all living things.

As long as an organism is alive, the amount of C-14 in its cellular structure remains constant. But when the organism dies, the amount of C-14 begins to decrease. Scientists know the half-life of C-14 (5,730 years, listed in Table 5-2), so they can figure out how long ago the organism died.

Radioactive dating using C-14 has been used to determine the age of skeletons found at archeological sites. Recently, it was used to date the *Shroud of Turin,* a piece of linen in the shape of a burial cloth that contains an image of a man. Many thought that it was the burial cloth of Jesus, but in 1988, radiocarbon dating determined that the cloth dated from around A.D. 1200–1300. Even though we don't know how the image of the man was placed on the Shroud, C-14 dating has proven that it's not the death cloth of Jesus.

Carbon-14 dating can only be used to determine the age of something that was once alive. It can't be used to determine the age of a moon rock or a meteorite. For nonliving substances, scientists use other isotopes, such as potassium-40.

Gone (Nuclear) Fission

In the 1930s, scientists discovered that some nuclear reactions can be initiated and controlled (see "Radioactivity and Man-Made Radioactive Decay," earlier in this chapter). Scientists usually accomplished this task by bombarding a large isotope with a second, smaller one — commonly a neutron.

The collision caused the larger isotope to break apart into two or more elements, which is called *nuclear fission.* The nuclear fission of uranium-235 is shown in the following equation:

$$^{235}_{92}U + ^{1}_{0}n \rightarrow ^{142}_{56}Ba + ^{91}_{36}Kr + 3\,^{1}_{0}n$$

Reactions of this type also release a lot of energy. Where does the energy come from? Well, if you make *very* accurate measurement of the masses of all the atoms and subatomic particles you start with and all the atoms and subatomic particles you end up with, and then compare the two, you find that there's some "missing" mass. Matter disappears during the nuclear reaction. This loss of matter is called the *mass defect.* The missing matter is converted into energy.

You can actually calculate the amount of energy produced during a nuclear reaction with a fairly simple equation developed by Einstein: $E = mc^2$. In this equation, E is the amount of energy produced, m is the "missing" mass, or the mass defect, and c is the speed of light, which is a rather large number. The speed of light is squared, making that part of the equation a *very* large number that, even when multiplied by a small amount of mass, yields a *large* amount of energy.

Chain reactions and critical mass

Take a look at the equation for the fission of U-235 in the preceding section. Notice that one neutron was used, but three were produced. These three neutrons, if they encounter other U-235 atoms, can initiate other fissions, producing even more neutrons. It's the old domino effect. In terms of nuclear chemistry, it's a continuing cascade of nuclear fissions called a *chain reaction.* The chain reaction of U-235 is shown in Figure 5-3.

This chain reaction depends on the release of more neutrons than were used during the nuclear reaction. If you were to write the equation for the nuclear fission of U-238, the more abundant isotope of uranium, you'd use one neutron and only get one back out. You can't have a chain reaction with U-238. But isotopes that produce an *excess* of neutrons in their fission support a chain reaction. This type of isotope is said to be *fissionable,* and there are only two main fissionable isotopes used during nuclear reactions — uranium-235 and plutonium-239.

A certain minimum amount of fissionable matter is needed to support a self-sustaining chain reaction, and it's related to those neutrons. If the sample is small, then the neutrons are likely to shoot out of the sample before hitting a U-235 nucleus. If they don't hit a U-235 nucleus, no extra electrons and no energy are released. The reaction just fizzles. The minimum amount of fissionable material needed to ensure that a chain reaction occurs is called the *critical mass.* Anything less than this amount is called *subcritical.*

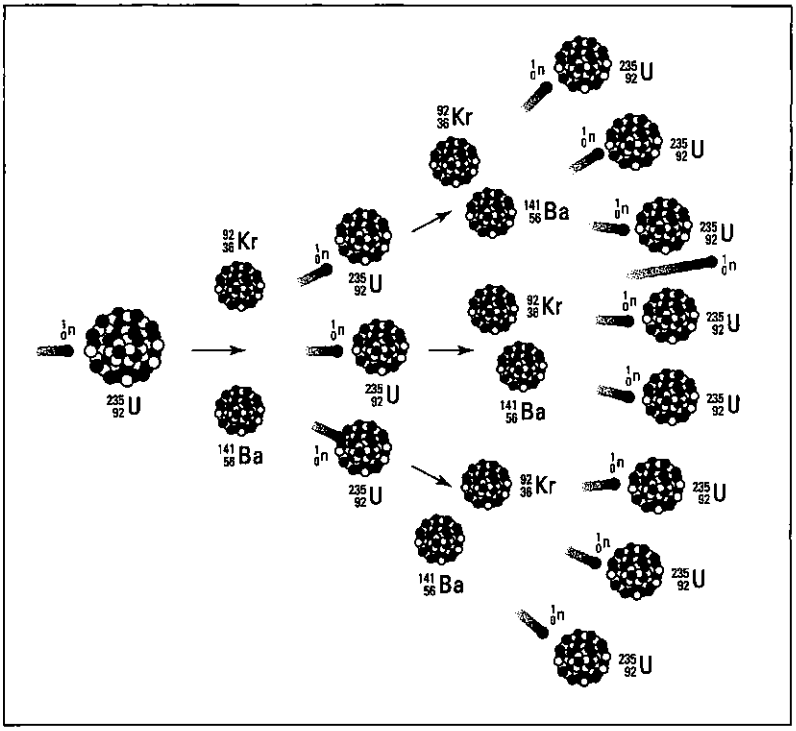


Figure 5-3: Chain reaction.

Atomic bombs (big bangs that aren't theories)

Because of the tremendous amount of energy released in a fission chain reaction, the military implications of nuclear reactions were immediately realized. The first atomic bomb was dropped on Hiroshima, Japan, on August 6, 1945.

In an atomic bomb, two pieces of a fissionable isotope are kept apart. Each piece, by itself, is subcritical. When it's time for the bomb to explode, conventional explosives force the two pieces together to cause a critical mass. The chain reaction is uncontrolled, releasing a tremendous amount of energy almost instantaneously.

The real trick, however, is to control the chain reaction, releasing its energy slowly so that ends other than destruction might be achieved.

Nuclear power plants

The secret to controlling a chain reaction is to control the neutrons. If the neutrons can be controlled, then the energy can be released in a controlled way. That's what scientists have done with nuclear power plants.

In many respects, a nuclear power plant is similar to a conventional fossil fuel power plant. In this type of plant, a fossil fuel (coal, oil, natural gas) is burned, and the heat is used to boil water, which, in turn, is used to make steam. The steam is then used to turn a turbine that is attached to a generator that produces electricity.

The big difference between a conventional power plant and a nuclear power plant is that the nuclear power plant produces heat through nuclear fission chain reactions.

How do nuclear power plants make electricity?

Most people believe that the concepts behind nuclear power plants are tremendously complex. That's really not the case. Nuclear power plants are very similar to conventional fossil fuel plants.

The fissionable isotope is contained in fuel rods in the reactor core. All the fuel rods together comprise the critical mass. Control rods, commonly made of boron or cadmium, are in the core, and they act like neutron sponges to control the rate of radioactive decay. Operators can stop a chain reaction completely by pushing the control rods all the way into the reactor core, where they absorb all the neutrons. The operators can then pull out the control rods a little at a time to produce the desired amount of heat.

A liquid (water or, sometimes, liquid sodium) is circulated through the reactor core, and the heat generated by the fission reaction is absorbed. The liquid then flows into a steam generator, where steam is produced as the heat is absorbed by water. This steam is then piped through a steam turbine that's connected to an electric generator. The steam is condensed and recycled through the steam generator. This forms a closed system; that is, no water or steam escapes — it's all recycled.

The liquid that circulates through the reactor core is also part of a closed system. This closed system helps ensure that no contamination of the air or water takes place. But sometimes problems do arise.

Oh, so many problems

In the United States, there are approximately 100 nuclear reactors, producing a little more than 20 percent of the country's electricity. In France, almost

80 percent of the country's electricity is generated through nuclear fission. Nuclear power plants have certain advantages. No fossil fuels are burned (saving fossil-fuel resources for producing plastics and medicines), and there are no combustion products, such as carbon dioxide, sulfur dioxide, and so on, to pollute the air and water. But problems are associated with nuclear power plants.

One is cost. Nuclear power plants are expensive to build and operate. The electricity that's generated by nuclear power costs about twice as much as electricity generated through fossil fuel or hydroelectric plants. Another problem is that the supply of fissionable uranium-235 is limited. Of all the naturally occurring uranium, only about 0.75 percent is U-235. A vast majority is nonfissionable U-238. At current usage levels, we'll be out of naturally occurring U-235 in fewer than 100 years. A little bit more time can be gained through the use of breeder reactors (see "Breeder reactors: Making more nuclear stuff," later in this chapter). But there's a limit to the amount of nuclear fuel available in the earth, just as there's a limit to the amount of fossil fuels.

However, the two major problems associated with nuclear fission power are accidents (safety) and disposal of nuclear wastes.

Accidents: Three Mile Island and Chernobyl

Although nuclear power reactors really do have a good safety record, the distrust and fear associated with radiation make most people sensitive to safety issues and accidents. The most serious accident to occur in the United States happened in 1979 at the Three Mile Island Plant in Pennsylvania. A combination of operator error and equipment failure caused a loss of reactor core coolant. The loss of coolant led to a partial meltdown and the release of a small amount of radioactive gas. There was no loss of life or injury to plant personnel or the general population.

This was not the case at Chernobyl, Ukraine, in 1986. Human error, along with poor reactor design and engineering, contributed to a tremendous overheating of the reactor core, causing it to rupture. Two explosions and a fire resulted, blowing apart the core and scattering nuclear material into the atmosphere. A small amount of this material made its way to Europe and Asia. The area around the plant is *still* uninhabitable. The reactor has been encased in concrete, and it must remain that way for hundreds of years. Hundreds of people died. Many others felt the effect of radiation poisoning. Instances of thyroid cancer, possibly caused by the release of I-13, have risen dramatically in the towns surrounding Chernobyl. It will be many more years until the effects of this disaster will be fully known.

How do you get rid of this stuff: Nuclear wastes

The fission process produces large amounts of radioactive isotopes. If you look at Table 5-2, you'll notice that some of the half-lives of radioactive isotopes are rather long. Those isotopes are safe after ten half-lives. The length of ten half-lives presents a problem when dealing with the waste products of a fission reactor.

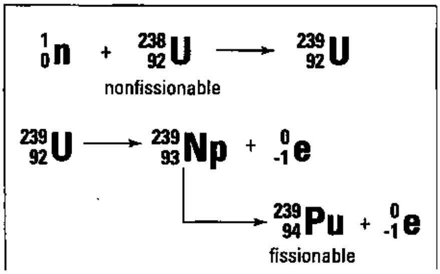
Eventually, all reactors must have their nuclear fuel replenished. And as we disarm nuclear weapons, we must deal with their radioactive material. Many of these waste products have long half-lives. How do we safely store the isotopes until their residual radioactivity has dropped to safe limits (ten half-lives)? How do we protect the environment and ourselves, and our children for generations to come, from this waste? These questions are undoubtedly the most serious problem associated with the peaceful use of nuclear power.

Nuclear waste is divided into low-level and high-level material, based on the amount of radioactivity being emitted. In the United States, low-level wastes are stored at the site of generation or at special storage facilities. The wastes are basically buried and guarded at the sites. High-level wastes pose a much larger problem. They're temporarily being stored at the site of generation, with plans to eventually seal the material in glass and then in drums. The material will then be stored underground in Nevada. At any rate, the waste must be kept safe and undisturbed for at least 10,000 years. Other countries face the same problems. There has been some dumping of nuclear material into deep trenches in the sea, but this practice has been discouraged by many nations.

Breeder reactors: Making more nuclear stuff

Only the U-235 isotope of uranium is fissionable because it's the only isotope of uranium that produces the excess of neutrons needed to maintain a chain-reaction. The far more plentiful U-238 isotope doesn't produce those extra neutrons.

The other commonly used fissionable isotope, plutonium-239 (Pu-239), is very rare in nature. But there's a way to make Pu-239 from U-238 in a special fission reactor called a *breeder reactor.* Uranium-238 is first bombarded with a neutron to produce U-239, which decays to Pu-239. The process is shown in Figure 5-4.



$$^{1}_{0}\mathrm{n} + \underset{\text{nonfissionable}}{^{238}_{92}\mathrm{U}} \longrightarrow {}^{239}_{92}\mathrm{U}$$

$$^{239}_{92}\mathrm{U} \longrightarrow {}^{239}_{93}\mathrm{Np} + {}^{0}_{-1}\mathrm{e}$$

$$^{239}_{93}\mathrm{Np} \longrightarrow \underset{\text{fissionable}}{^{239}_{94}\mathrm{Pu}} + {}^{0}_{-1}\mathrm{e}$$

Figure 5-4: The breeder reactor process.

Breeder reactors can extend the supply of fissionable fuels for many, many years, and they're currently being used in France. But the United States is moving *slowly* with the construction of breeder reactors because of several problems associated with them. First, they're extremely expensive to build. Second, they produce large amounts of nuclear wastes. And, finally, the plutonium that's produced is much more hazardous to handle than uranium and can easily be used in an atomic bomb.

Nuclear Fusion: The Hope for Our Energy Future

Soon after the fission process was discovered, another process, called *fusion,* was discovered. Fusion is essentially the opposite of fission. In fission, a heavy nucleus is split into smaller nuclei. With fusion, lighter nuclei are fused into a heavier nucleus.

The fusion process is the reaction that powers the sun. On the sun, in a series of nuclear reactions, four isotopes of hydrogen-1 are fused into a helium-4 with the release of a tremendous amount of energy. Here on earth, two other isotopes of hydrogen are used: H-2, called deuterium, and H-3, called tritium. Deuterium is a minor isotope of hydrogen, but it's still relatively abundant. Tritium doesn't occur naturally, but it can easily be produced by bombarding deuterium with a neutron. The fusion reaction is shown in the following equation:

$$^{3}_{1}H + ^{2}_{1}H \rightarrow ^{4}_{2}He + ^{1}_{0}n$$

The first demonstration of nuclear fusion — the hydrogen bomb — was conducted by the military. A hydrogen bomb is approximately 1,000 times as powerful as an ordinary atomic bomb.

The isotopes of hydrogen needed for the hydrogen bomb fusion reaction were placed around an ordinary fission bomb. The explosion of the fission bomb released the energy needed to provide the *activation energy* (the energy necessary to initiate, or start, the reaction) for the fusion process.

Control issues

The goal of scientists for the last 45 years has been the controlled release of energy from a fusion reaction. If the energy from a fusion reaction can be released slowly, it can be used to produce electricity. It will provide an unlimited supply of energy that has no wastes to deal with or contaminants to harm the atmosphere — simply non-polluting helium. But achieving this goal requires overcoming three problems:

- Temperature
- Time
- Containment

Temperature

The fusion process requires an extremely high activation energy. Heat is used to provide the energy, but it takes a *lot* of heat to start the reaction. Scientists estimate that the sample of hydrogen isotopes must be heated to approximately 40,000,000 K. (K represents the Kelvin temperature scale. To get the Kelvin temperature, you add 273 to the Celsius temperature. Chapter 2 explains all about Kelvin and his pals Celsius and Fahrenheit.)

Now 40,000,000 K is hotter than the sun! At this temperature, the electrons have long since left the building; all that's left is a positively-charged *plasma,* bare nuclei heated to a tremendously high temperature. Presently, scientists are trying to heat samples to this high temperature through two ways — magnetic fields and lasers. Neither one has yet achieved the necessary temperature.

Time

Time is the second problem scientists must overcome to achieve the controlled release of energy from fusion reactions. The charged nuclei must be held together close enough and long enough for the fusion reaction to start. Scientists estimate that the plasma needs to be held together at 40,000,000 K for about one second.

Containment

Containment is the major problem facing fusion research. At 40,000,000 K, everything is a gas. The best ceramics developed for the space program would vaporize when exposed to this temperature. Because the plasma has a charge, magnetic fields can be used to contain it — like a magnetic bottle. But if the bottle leaks, the reaction won't take place. And scientists have yet to create a magnetic field that won't allow the plasma to leak. Using lasers to zap the hydrogen isotope mixture and provide the necessary energy bypasses the containment problem. But scientists have not figured out how to protect the lasers themselves from the fusion reaction.

What the future holds

The latest estimates indicate that science is 5 to 10 years away from showing that fusion can work: This is the so-called *break-even point,* where we get out more energy than we put in. It will then be another 20 to 30 years before a functioning fusion reactor is developed. But scientists are optimistic that controlled fusion power will be achieved. The rewards are great — an unlimited source of nonpolluting energy.

An interesting by-product of fusion research is the *fusion torch* concept. With this idea, the fusion plasma, which must be cooled in order to produce steam, is used to incinerate garbage and solid wastes. Then the individual atoms and small molecules that are produced are collected and used as raw materials for industry. It seems like an ideal way to close the loop between waste and raw materials. Time will tell if this concept will eventually make it into practice.

Am I Glowing? The Effects of Radiation

Radiation can have two basic effects on the body:

- It can destroy cells with heat.
- It can ionize and fragment cells.

Radiation generates heat. This heat can destroy tissue, much like a sunburn does. In fact, the term *radiation burn* is commonly used to describe the destruction of skin and tissue due to heat.

The other major way that radiation can affect the body is through the ionization and fragmentation of cells. Radioactive particles and radiation have a lot of *kinetic energy* (energy of motion — see Chapter 2) associated with them. When these particles strike cells within the body, they can *fragment* (destroy) the cells or *ionize* the cells — turn the cells into ions (charged atoms) by knocking off an electron. (Flip to Chapter 3 for the full scoop on ions.) Ionization weakens bonds and can lead to the damage, destruction, or mutation of the cells.

Radon: Hiding in our houses

Radon is a radioactive isotope that's been receiving a lot of publicity recently. Radon-222 is formed naturally as part of the decay of uranium. It's an unreactive noble gas, so it escapes from the ground into the air. Because it's heavier than air, it can accumulate in basements.

Radon itself has a short half-life of 3.8 days, but it decays to Polonium-218, a solid. So if radon is inhaled, solid Po-218 can accumulate in the lungs. Po-218 is an alpha emitter, and, even though this type of radiation is not very penetrating, it has been linked to increased instances of lung cancer. In many parts of the United States, radon testing is performed before selling a house. Commercial test kits can be opened, left in the basement area for a specified amount of time, and then sent to a lab for analysis. The question of whether radon represents a serious problem is still being investigated and debated.