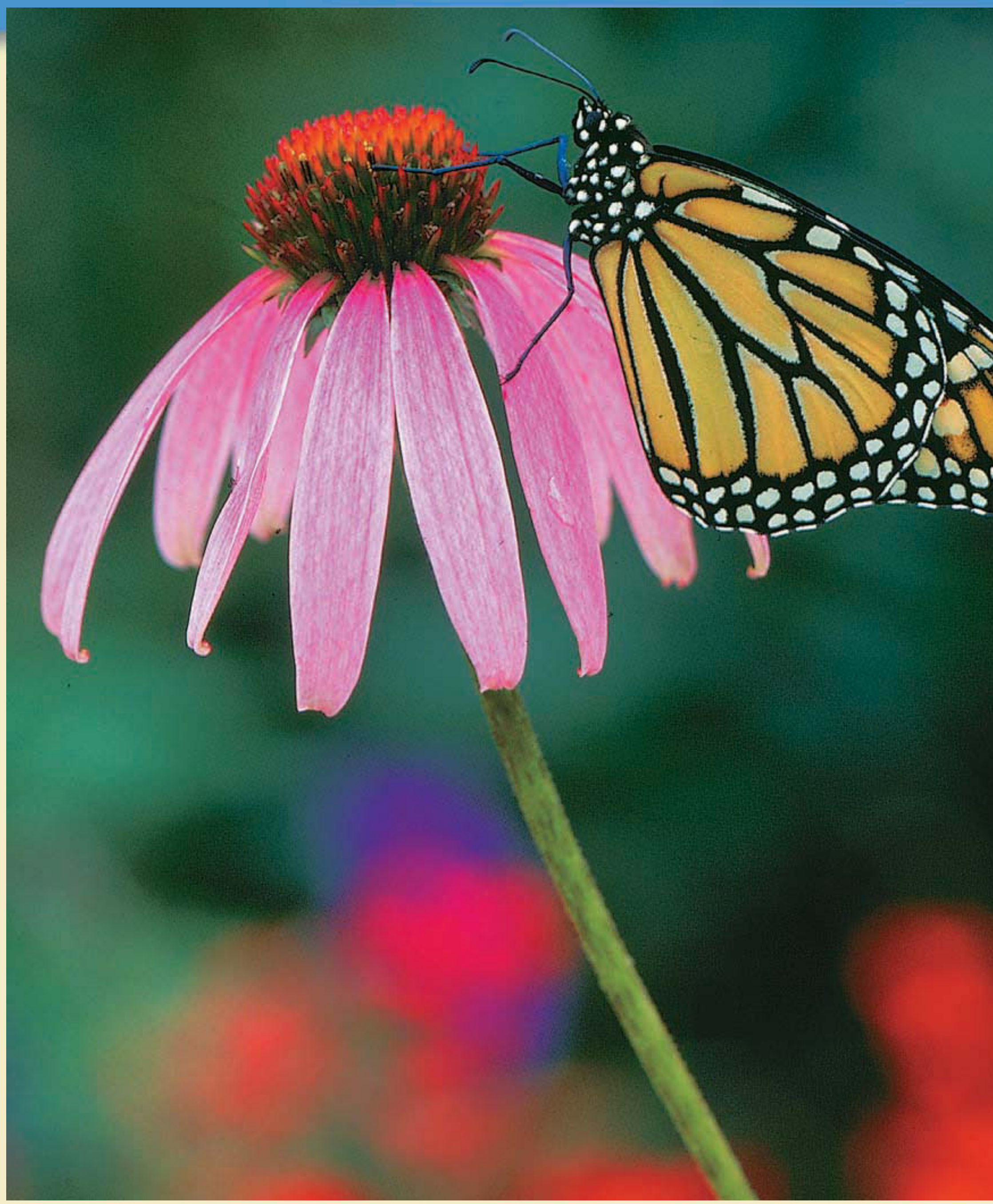
1 Chemical Foundations

Contents



Male Monarch butterflies use the pheromones produced by a gland on their wings to make themselves attractive to females.

When you start your car, do you think about chemistry? Probably not, but you should. The power to start your car is furnished by a lead storage battery. How does this battery work, and what does it contain? When a battery goes dead, what does that mean? If you use a friend's car to "jump start" your car, did you know that your battery could explode? How can you avoid such an unpleasant possibility? What is in the gasoline that you put in your tank, and how does it furnish the energy to drive to school? What is the vapor that comes out of the exhaust pipe, and why does it cause air pollution? Your car's air conditioner might have a substance in it that is leading to the destruction of the ozone layer in the upper atmosphere. What are we doing about that? And why is the ozone layer important anyway?

All these questions can be answered by understanding some chemistry. In fact, we'll consider the answers to all these questions in this text.

Chemistry is around you all the time. You are able to read and understand this sentence because chemical reactions are occurring in your brain. The food you ate for breakfast or lunch is now furnishing energy through chemical reactions. Trees and grass grow because of chemical changes.

Chemistry also crops up in some unexpected places. When archaeologist Luis Alvarez was studying in college, he probably didn't realize that the chemical elements iridium and niobium would make him very famous when they helped him solve the problem of the disappearing dinosaurs. For decades scientists had wrestled with the mystery of why the dinosaurs, after ruling the earth for millions of years, suddenly became extinct 65 million years ago. In studying core samples of rocks dating back to that period, Alvarez and his coworkers recognized unusual levels of iridium and niobium in these samples—levels much more characteristic of extraterrestrial bodies than of the earth. Based on these observations, Alvarez hypothesized that a large meteor hit the earth 65 million years ago, changing atmospheric conditions so much that the dinosaurs' food couldn't grow, and they died—almost instantly in the geologic timeframe.

Chemistry is also important to historians. Did you realize that lead poisoning probably was a significant contributing factor to the decline of the Roman Empire? The Romans had high exposure to lead from lead-glazed pottery, lead water pipes, and a sweetening syrup called *sapa* that was prepared by boiling down grape juice in lead-lined vessels. It turns out that one reason for sapa's sweetness was lead acetate ("sugar of lead") that formed as the juice was cooked down. Lead poisoning with its symptoms of lethargy and mental malfunctions certainly could have contributed to the demise of the Roman society.

Chemistry is also apparently very important in determining a person's behavior. Various studies have shown that many personality disorders can be linked directly to imbalances of trace elements in the body. For example, studies on the inmates at Stateville Prison in Illinois have linked low cobalt levels with violent behavior. Lithium salts have been shown to be very effective in controlling the effects of manic depressive disease, and you've probably at some time in your life felt a special "chemistry" for another person. Studies suggest there is literally chemistry going on between two people who are attracted to each other. "Falling in love" apparently causes changes in the chemistry of the brain; chemicals are produced that give that "high" associated with a new relationship. Unfortunately, these chemical effects seem to wear off over time, even if the relationship persists and grows.

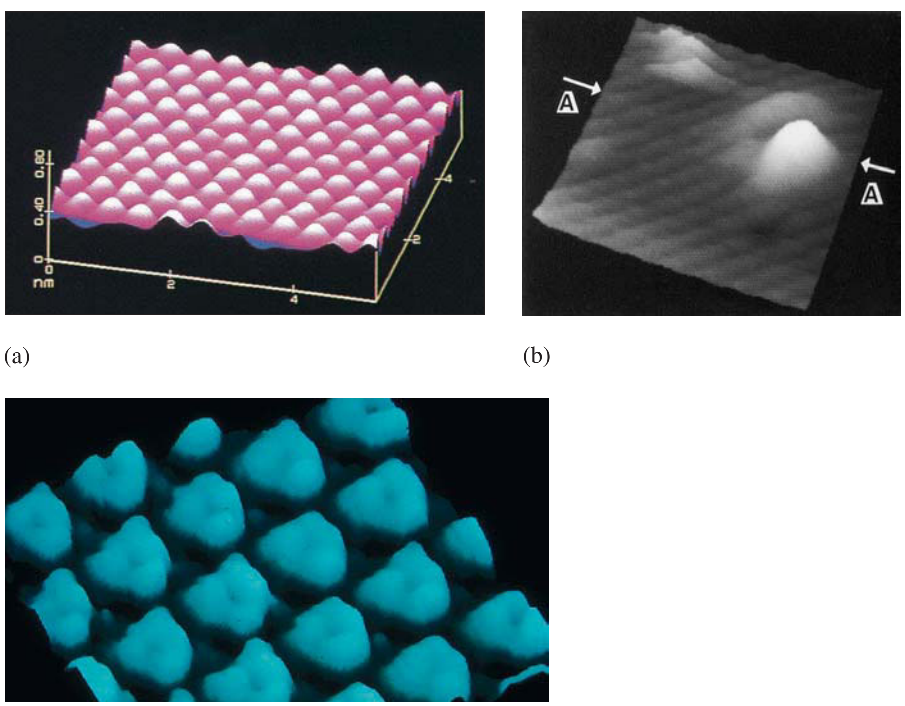
The importance of chemistry in the interactions of people should not really surprise us, since we know that insects communicate by emitting and receiving chemical signals via molecules called *pheromones*. For example, ants have a very complicated set of chemical signals to signify food sources, danger, and so forth. Also, various female sex attractants have been isolated and used to lure males into traps to control insect populations. It would not be surprising if humans also emitted chemical signals that we were not aware of on a conscious level. Thus chemistry is pretty interesting and pretty important. The main goal of this text is to help you understand the concepts of chemistry so that you can better appreciate the world around you and can be more effective in whatever career you choose.

1.1 Chemistry: An Overview

Since the time of the ancient Greeks, people have wondered about the answer to the question: What is matter made of? For a long time humans have believed that matter is composed of atoms, and in the previous three centuries we have collected much indirect evidence to support this belief. Very recently, something exciting has happened—for the first time we can "see" individual atoms. Of course, we cannot see atoms with the naked eye but must use a special microscope called a scanning tunneling microscope (STM). Although we will not consider the details of its operation here, the STM uses an electron current from a tiny needle to probe the surface of a substance. The STM pictures of several substances are shown in Fig. 1.1. Notice how the atoms are connected to one another by "bridges," which, as we will see, represent the electrons that interconnect atoms.

In addition to "seeing" the atoms in solids such as salt, we have learned how to isolate and view a single atom. For example, the tiny white dot in the center of Fig. 1.2 is a single mercury atom that is held in a special trap.

So, at this point, we are fairly sure that matter consists of individual atoms. The nature of these atoms is quite complex, and the components of atoms don't behave much like the objects we see in the world of our experience. We call this world the *macroscopic world*—the world of cars, tables, baseballs, rocks, oceans, and so forth. One of the main jobs of a scientist is to delve into the macroscopic world and discover its "parts." For example, when you view a beach from a distance, it looks like a continuous solid substance. As you get closer, you see that the beach is really made up of individual grains of sand.



(a) (b) (c)

FIGURE 1.1
(a) The surface of a single grain of table salt. (b) An oxygen atom (indicated by arrow) on a gallium arsenide surface. (c) Scanning tunneling microscope image showing rows of ring-shaped clusters of benzene molecules on a rhodium surface. Each "doughnut"-shaped image represents a benzene molecule.

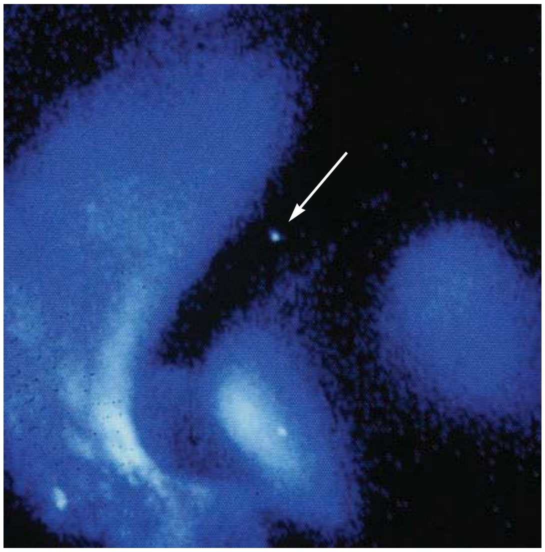
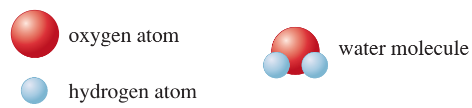


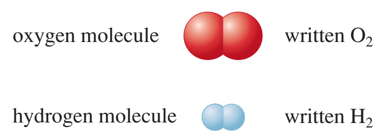
FIGURE 1.2
A charged mercury atom shows up as a tiny white dot (indicated by the arrow).

As we examine these grains of sand, we find they are composed of silicon and oxygen atoms connected to each other to form intricate shapes (see Fig. 1.3). One of the main challenges of chemistry is to understand the connection between the macroscopic world that we experience and the *microscopic world* of atoms and molecules. To truly understand chemistry you must learn to think on the atomic level. We will spend much time in this text helping you learn to do that.

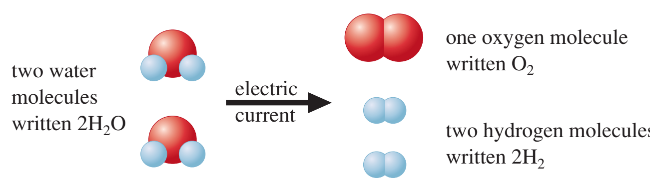
One of the amazing things about our universe is that the tremendous variety of substances we find there results from only about 100 different kinds of atoms. You can think of these approximately 100 atoms as the letters in an alphabet out of which all the "words" in the universe are made. It is the way the atoms are organized in a given substance that determines the properties of that substance. For example, water, one of the most common and important substances on earth, is composed of two types of atoms: hydrogen and oxygen. There are two hydrogen atoms and one oxygen atom bound together to form the water molecule:



When an electric current passes through it, water is decomposed to hydrogen and oxygen. These *chemical elements* themselves exist naturally as diatomic (two-atom) molecules:



We can represent the decomposition of water to its component elements, hydrogen and oxygen, as follows:



Notice that it takes two molecules of water to furnish the right number of oxygen and hydrogen atoms to allow for the formation of the two-atom molecules. This reaction explains

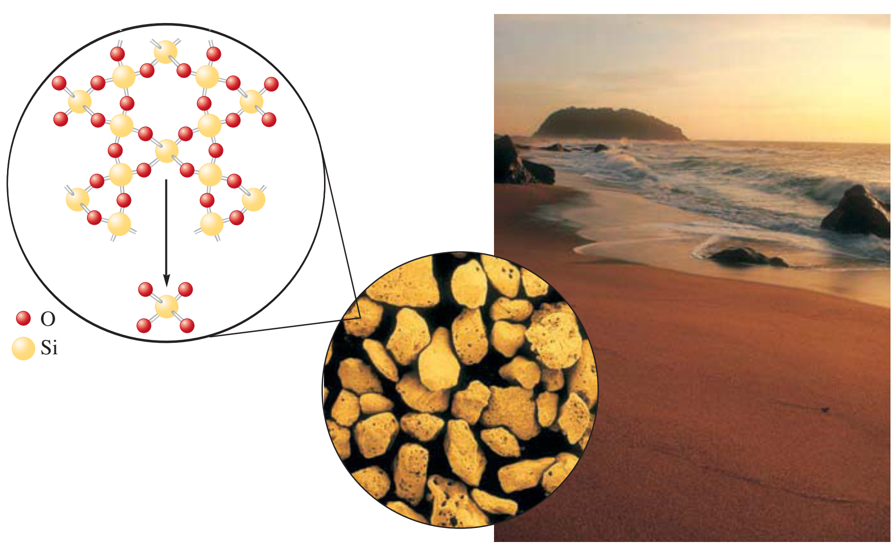


FIGURE 1.3
Sand on a beach looks uniform from a distance, but up close the irregular sand grains are visible, and each grain is composed of tiny atoms.

CHEMICAL IMPACT

The Chemistry of Art

The importance of chemistry can show up in some unusual places. For example, a knowledge of chemistry is crucial to authenticating, preserving, and restoring art objects. The J. Paul Getty Museum in Los Angeles has a state-of-the-art chemical laboratory that costs many millions of dollars and employs many scientists. The National Gallery of Art (NGA) in Washington, D.C., also operates a highly sophisticated laboratory that employs 10 people: five chemists, a botanist, an art historian, a technician with a chemistry degree, and two fellows (interns).

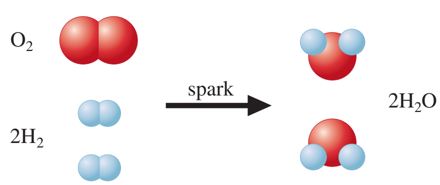
One of the chemists at NGA is Barbara Berrie, who specializes in identifying paint pigments. One of her duties is to analyze a painting to see whether the paint pigments are appropriate for the time the picture was supposedly painted and consistent with the pigments known to be used by the artist given credit for the painting. This analysis is one way in which paintings can be authenticated. One of Berrie's recent projects was to analyze the 1617 oil painting *St. Cecilia and an Angel*. Her results showed the painting was the work of two artists of the time, Orazio Gentileschi and Giovanni Lanfranco. Originally the work was thought to be by Gentileschi alone.

Berrie is also working to define the range of colors used by water colorist Winslow Homer (the NGA has 30 Homer paintings in its collection) and to show how his color palette changed over his career. In addition, she is exploring how acidity affects the decomposition of a particular deep green transparent pigment (called copper resinate) used by Italian Renaissance artists so that paintings using this pigment can be better preserved.

Berrie says, "The chemistry I do is not hot-dog chemistry, just good old-fashioned general chemistry."

Dr. Barbara Berrie of the National Gallery of Art is shown analyzing the glue used in the wooden supports for a 14th century altar piece.

why the battery in your car can explode if you jump start it improperly. When you hook up the jumper cables, current flows through the dead battery, which contains water (and other things), and causes hydrogen and oxygen to form by decomposition of some of the water. A spark can cause this accumulated hydrogen and oxygen to explode, forming water again.



This example illustrates two of the fundamental concepts of chemistry: (1) matter is composed of various types of atoms, and (2) one substance changes to another by reorganizing the way the atoms are attached to each other.

These are core ideas of chemistry, and we will have much more to say about them.

Science: A Process for Understanding Nature and Its Changes

How do you tackle the problems that confront you in real life? Think about your trip to school. If you live in a city, traffic is undoubtedly a problem you confront daily. How do you decide the best way to drive to school? If you are new in town, you first get a map and look at the possible ways to make the trip. Then you might collect information from people who know the area about the advantages and disadvantages of various routes. Based on this information, you probably try to predict the best route. However, you can find the best route only by trying several of them and comparing the results. After a few experiments with the various possibilities, you probably will be able to select the best way. What you are doing in solving this everyday problem is applying the same process that scientists use to study nature. The first thing you did was collect relevant data. Then you made a prediction, and then you tested it by trying it out. This process contains the fundamental elements of science.

1. Making observations (collecting data)
2. Making a prediction (formulating a hypothesis)
3. Doing experiments to test the prediction (testing the hypothesis)

Scientists call this process the *scientific method*. We will discuss it in more detail in the next section. One of life's most important activities is solving problems—not "plug and chug" exercises, but real problems—problems that have new facets to them, that involve things you may have never confronted before. The more creative you are at solving these problems, the more effective you will be in your career and your personal life. Part of the reason for learning chemistry, therefore, is to become a better problem solver. Chemists are usually excellent problem solvers, because to master chemistry, you have to master the scientific approach. Chemical problems are frequently very complicated—there is usually no neat and tidy solution. Often it is difficult to know where to begin.

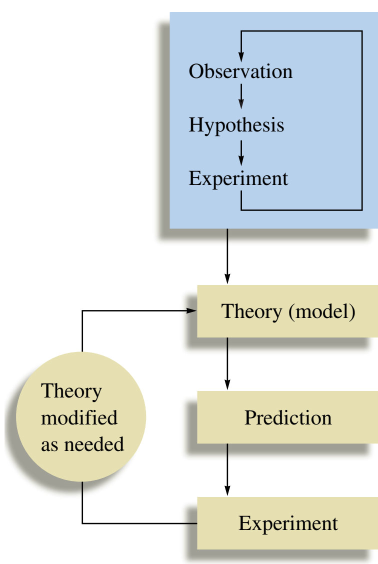


FIGURE 1.4
The fundamental steps of the scientific method.

1.2 The Scientific Method

Science is a framework for gaining and organizing knowledge. Science is not simply a set of facts but also a plan of action—a *procedure* for processing and understanding certain types of information. Scientific thinking is useful in all aspects of life, but in this text we will use it to understand how the chemical world operates. As we have said in our previous discussion, the process that lies at the center of scientific inquiry is called the **scientific method.** There are actually many scientific methods, depending on the nature of the specific problem under study and on the particular investigator involved. However, it is useful to consider the following general framework for a generic scientific method (see Fig. 1.4):

Steps in the Scientific Method

➡ 1 **_Making observations._ Observations may be _qualitative_ (the sky is blue; water is a liquid) or _quantitative_ (water boils at 100°C; a certain chemistry book weighs 2 kilograms). A qualitative observation does not involve a number. A quantitative observation (called a measurement) involves both a number and a unit.**

➡ 2 **_Formulating hypotheses._ A hypothesis is a _possible_ explanation for an observation.**

➡ 3 **_Performing experiments._ An experiment is carried out to test a hypothesis. This involves gathering new information that enables a scientist to decide whether**

the hypothesis is valid—that is, whether it is supported by the new information learned from the experiment. Experiments always produce new observations, and this brings the process back to the beginning again.

To understand a given phenomenon, these steps are repeated many times, gradually accumulating the knowledge necessary to provide a possible explanation of the phenomenon.

Scientific Models

Once a set of hypotheses that agrees with the various observations is obtained, the hypotheses are assembled into a theory. A **theory,** which is often called a **model,** is a set of tested hypotheses that gives an overall explanation of some natural phenomenon.

It is very important to distinguish between observations and theories. An observation is something that is witnessed and can be recorded. A theory is an *interpretation*—a possible explanation of *why* nature behaves in a particular way. Theories inevitably change as more information becomes available. For example, the motions of the sun and stars have remained virtually the same over the thousands of years during which humans have been observing them, but our explanations—our theories—for these motions have changed greatly since ancient times. (See the Chemical Impact on Observations, Theories, and the Planets on the Web site.)

The point is that scientists do not stop asking questions just because a given theory seems to account satisfactorily for some aspect of natural behavior. They continue doing experiments to refine or replace the existing theories. This is generally done by using the currently accepted theory to make a prediction and then performing an experiment (making a new observation) to see whether the results bear out this prediction.

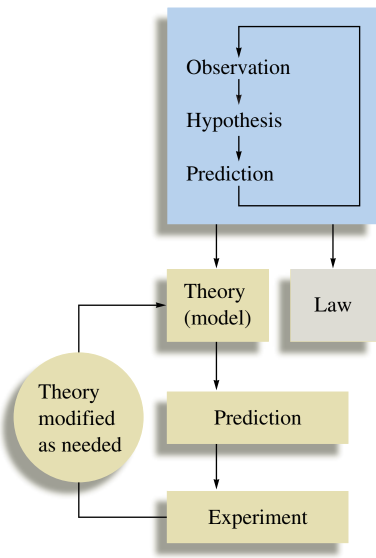


FIGURE 1.5
The various parts of the scientific method.

Always remember that theories (models) are human inventions. They represent attempts to explain observed natural behavior in terms of human experiences. A theory is actually an educated guess. We must continue to do experiments and to refine our theories (making them consistent with new knowledge) if we hope to approach a more nearly complete understanding of nature.

As scientists observe nature, they often see that the same observation applies to many different systems. For example, studies of innumerable chemical changes have shown that the total observed mass of the materials involved is the same before and after the change. Such generally observed behavior is formulated into a statement called a **natural law.** For example, the observation that the total mass of materials is not affected by a chemical change in those materials is called the **law of conservation of mass.**

Note the difference between a natural law and a theory. A natural law is a summary of observed (measurable) behavior, whereas a theory is an explanation of behavior. *A law summarizes what happens; a theory (model) is an attempt to explain why it happens.*

In this section we have described the scientific method as it might ideally be applied (see Fig. 1.5). However, it is important to remember that science does not always progress smoothly and efficiently. For one thing, hypotheses and observations are not totally independent of each other, as we have assumed in the description of the idealized scientific

Robert Boyle (1627–1691) was born in Ireland. He became especially interested in experiments involving air and developed an air pump with which he produced evacuated cylinders. He used these cylinders to show that a feather and a lump of lead fall at the same rate in the absence of air resistance and that sound cannot be produced in a vacuum. His most famous experiments involved careful measurements of the volume of a gas as a function of pressure. In his book *The Skeptical Chymist,* Boyle urged that the ancient view of elements as mystical substances should be abandoned and that an element should instead be defined as anything that cannot be broken down into simpler substances. This conception was an important step in the development of modern chemistry.

CHEMICAL IMPACT

A Note-able Achievement

Post-it Notes, a product of the 3M Corporation, revolutionized casual written communications and personal reminders. Introduced in the United States in 1980, these sticky-but-not-too-sticky notes have now found countless uses in offices, cars, and homes throughout the world.

The invention of sticky notes occurred over a period of about 10 years and involved a great deal of serendipity. The adhesive for Post-it Notes was discovered by Dr. Spencer F. Silver of 3M in 1968. Silver found that when an acrylate polymer material was made in a particular way, it formed cross-linked microspheres. When suspended in a solvent and sprayed on a sheet of paper, this substance formed a "sparse monolayer" of adhesive after the solvent evaporated. Scanning electron microscope images of the adhesive show that it has an irregular surface, a little like the surface of a gravel road. In contrast, the adhesive on cellophane tape looks smooth and uniform, like a superhighway. The bumpy surface of Silver's adhesive caused it to be sticky but not so sticky to produce permanent adhesion, because the number of contact points between the binding surfaces was limited.

When he invented this adhesive, Silver had no specific ideas for its use, so he spread the word of his discovery to his fellow employees at 3M to see if anyone had an application for it. In addition, over the next several years development was carried out to improve the adhesive's properties. It was not until 1974 that the idea for Post-it Notes popped up. One Sunday Art Fry, a chemical engineer for 3M, was singing in his church choir when he became annoyed that the bookmark in his hymnal kept falling out. He thought to himself that it would be nice if the bookmark were sticky enough to stay in place but not so sticky that it couldn't be moved. Luckily, he remembered Silver's glue—and the Post-it Note was born.

For the next three years Fry worked to overcome the manufacturing obstacles associated with the product. By 1977 enough Post-it Notes were being produced to supply 3M's corporate headquarters, where the employees quickly became addicted to their many uses. Post-it Notes are now available in 62 colors and 25 shapes.

In the years since their introduction, 3M has heard some remarkable stories connected to the use of these notes. For example, a Post-it Note was applied to the nose of a corporate jet, where it was intended to be read by the plane's Las Vegas ground crew. Someone forgot to remove it, however. The note was still on the nose of the plane when it landed in Minneapolis, having survived a take-off and landing and speeds of 500 miles per hour at temperatures as low as $-56°F$. Stories on the 3M Web site also describe how a Post-it Note on the front door of a home survived the 140 mile per hour winds of Hurricane Hugo and how a foreign official accepted Post-it Notes in lieu of cash when a small bribe was needed to cut through bureaucratic hassles.

Post-it Notes have definitely changed the way we communicate and remember things.

method. The coupling of observations and hypotheses occurs because once we begin to proceed down a given theoretical path, our hypotheses are unavoidably couched in the language of that theory. In other words, we tend to see what we expect to see and often fail to notice things that we do not expect. Thus the theory we are testing helps us because it focuses our questions. However, at the very same time, this focusing process may limit our ability to see other possible explanations.

It is also important to keep in mind that scientists are human. They have prejudices; they misinterpret data; they become emotionally attached to their theories and thus lose objectivity; and they play politics. Science is affected by profit motives, budgets, fads, wars, and religious beliefs. Galileo, for example, was forced to recant his astronomical observations in the face of strong religious resistance. Lavoisier, the father of modern chemistry, was beheaded because of his political affiliations. Great progress in the chemistry of nitrogen fertilizers resulted from the desire to produce explosives to fight wars. The progress of science is often affected more by the frailties of humans and their institutions than by the limitations of scientific measuring devices. The scientific methods are only as effective as the humans using them. They do not automatically lead to progress.

CHEMICAL IMPACT

Critical Units!

How important are conversions from one unit to another? If you ask the National Aeronautics and Space Administration (NASA), very important! In 1999 NASA lost a $125 million Mars Climate Orbiter because of a failure to convert from English to metric units.

The problem arose because two teams working on the Mars mission were using different sets of units. NASA's scientists at the Jet Propulsion Laboratory in Pasadena, California, assumed that the thrust data for the rockets on the Orbiter they received from Lockheed Martin Astronautics in Denver, which built the spacecraft, were in metric units. In reality, the units were English. As a result the Orbiter dipped 100 kilometers lower into the Mars atmosphere than planned and the friction from the atmosphere caused the craft to burn up.

NASA's mistake refueled the controversy over whether Congress should require the United States to switch to the metric system. About 95% of the world now uses the metric system, and the United States is slowly switching from English to metric. For example, the automobile industry has adopted metric fasteners and we buy our soda in two-liter bottles.

Units can be very important. In fact, they can mean the difference between life and death on some occasions. In 1983, for example, a Canadian jetliner almost ran out of fuel when someone pumped 22,300 pounds of fuel into the aircraft instead of 22,300 kilograms. Remember to watch your units!

Artist's conception of the lost Mars Climate Orbiter.

Soda is commonly sold in 2-liter bottles—an example of the use of SI units in everyday life.

1.3 Units of Measurement

Making observations is fundamental to all science. A quantitative observation, or *measurement,* always consists of two parts: a *number* and a scale (called a *unit*). Both parts must be present for the measurement to be meaningful.

In this textbook we will use measurements of mass, length, time, temperature, electric current, and the amount of a substance, among others. Scientists recognized long ago that standard systems of units had to be adopted if measurements were to be useful. If every scientist had a different set of units, complete chaos would result. Unfortunately, different standards were adopted in different parts of the world. The two major systems are the *English system* used in the United States and the *metric system* used by most of the rest of the industrialized world. This duality causes a good deal of trouble; for example, parts as simple as bolts are not interchangeable between machines built using the two systems. As a result, the United States has begun to adopt the metric system.

Most scientists in all countries have for many years used the metric system. In 1960, an international agreement set up a system of units called the *International System* (*le Système International* in French), or the **SI system.** This system is based on the metric system and units derived from the metric system. The fundamental SI units are listed in Table 1.1. We will discuss how to manipulate these units later in this chapter.

Because the fundamental units are not always convenient (expressing the mass of a pin in kilograms is awkward), prefixes are used to change the size of the unit. These are listed in Table 1.2. Some common objects and their measurements in SI units are listed in Table 1.3.

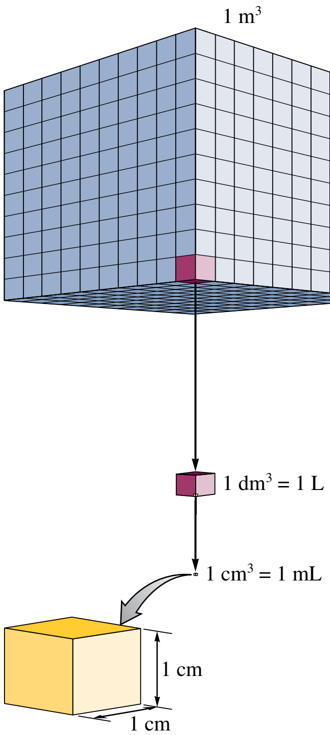


FIGURE 1.6
The largest cube has sides 1 m in length and a volume of 1 m^3. The middle-sized cube has sides 1 dm in length and a volume of 1 dm^3, or 1 L. The smallest cube has sides 1 cm in length and a volume of 1 cm^3, or 1 mL.

TABLE 1.1 The Fundamental SI Units

Physical Quantity	Name of Unit	Abbreviation
Mass	kilogram	kg
Length	meter	m
Time	second	s
Temperature	kelvin	K
Electric current	ampere	A
Amount of substance	mole	mol
Luminous intensity	candela	cd

One physical quantity that is very important in chemistry is *volume,* which is not a fundamental SI unit but is derived from length. A cube that measures 1 meter (m) on each edge is represented in Fig. 1.6. This cube has a volume of $(1\ \text{m})^3 = 1\ \text{m}^3$. Recognizing that there are 10 decimeters (dm) in a meter, the volume of this cube is $(1\ \text{m})^3 = (10\ \text{dm})^3 = 1000\ \text{dm}^3$. A cubic decimeter, that is $(1\ \text{dm})^3$, is commonly called a *liter (L),* which is a unit of volume slightly larger than a quart. As shown in Fig. 1.6, 1000 liters are contained in a cube with a volume of 1 cubic meter. Similarly, since 1 decimeter equals 10 centimeters (cm), the liter can be divided into 1000 cubes each with a volume of 1 cubic centimeter:

$$1\ \text{liter} = (1\ \text{dm})^3 = (10\ \text{cm})^3 = 1000\ \text{cm}^3$$

Also, since $1\ \text{cm}^3 = 1$ milliliter (mL),

$$1\ \text{liter} = 1000\ \text{cm}^3 = 1000\ \text{mL}$$

Thus 1 liter contains 1000 cubic centimeters, or 1000 milliliters.

Chemical laboratory work frequently requires measurement of the volumes of liquids. Several devices for the accurate determination of liquid volume are shown in Fig. 1.7.

An important point concerning measurements is the relationship between mass and weight. Although these terms are sometimes used interchangeably, they are *not* the same.

TABLE 1.2 The Prefixes Used in the SI System (Those most commonly encountered are shown in blue.)

Prefix	Symbol	Meaning	Exponential Notation*
exa	E	1,000,000,000,000,000,000	10^{18}
peta	P	1,000,000,000,000,000	10^{15}
tera	T	1,000,000,000,000	10^{12}
giga	G	1,000,000,000	10^{9}
mega	M	1,000,000	10^{6}
kilo	k	1,000	10^{3}
hecto	h	100	10^{2}
deka	da	10	10^{1}
—	—	1	10^{0}
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}
pico	p	0.000000000001	10^{-12}
femto	f	0.000000000000001	10^{-15}
atto	a	0.000000000000000001	10^{-18}

*See Appendix 1.1 if you need a review of exponential notation.

TABLE 1.3 Some Examples of Commonly Used Units

Length	A dime is 1 mm thick. A quarter is 2.5 cm in diameter. The average height of an adult man is 1.8 m.
Mass	A nickel has a mass of about 5 g. A 120-lb person has a mass of about 55 kg.
Volume	A 12-oz can of soda has a volume of about 360 mL.

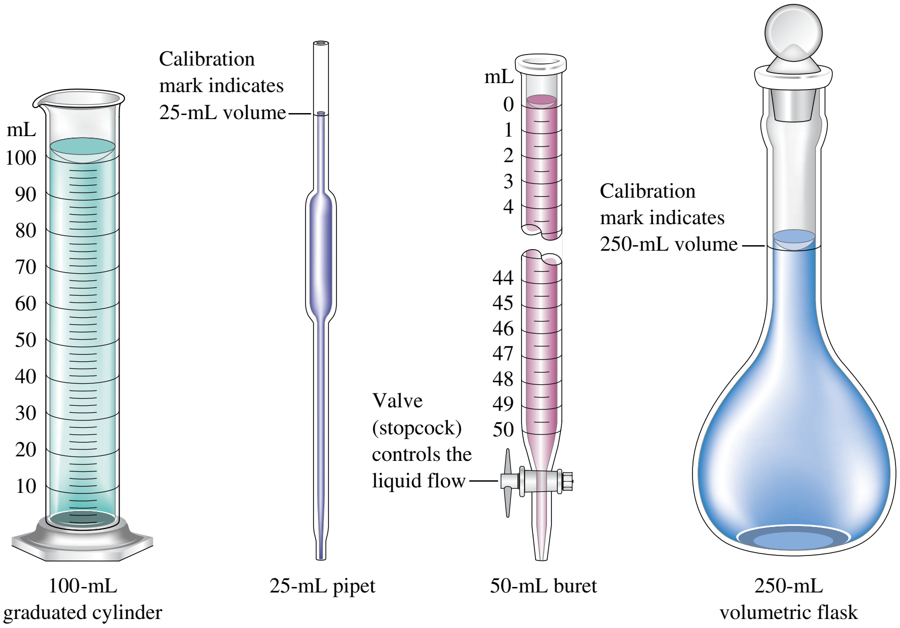


FIGURE 1.7
Common types of laboratory equipment used to measure liquid volume.

FIGURE 1.8
An electronic analytical balance.

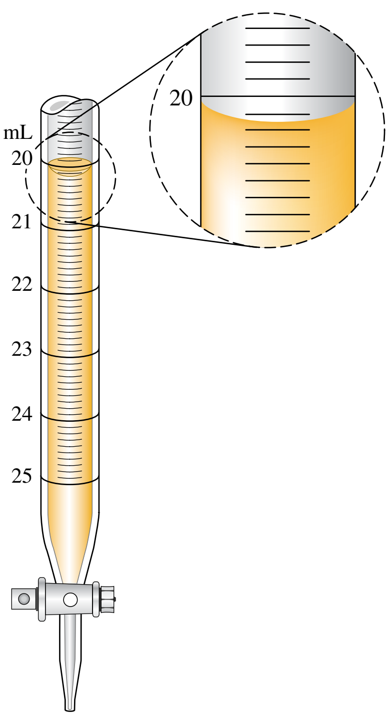


FIGURE 1.9
Measurement of volume using a buret. The volume is read at the bottom of the liquid curve (called the meniscus).

Mass *is a measure of the resistance of an object to a change in its state of motion*. Mass is measured by the force necessary to give an object a certain acceleration. On earth we use the force that gravity exerts on an object to measure its mass. We call this force the object's **weight.** Since weight is the response of mass to gravity, it varies with the strength of the gravitational field. Therefore, your body mass is the same on the earth or on the moon, but your weight would be much less on the moon than on earth because of the moon's smaller gravitational field.

Because weighing something on a chemical balance (see Fig. 1.8) involves comparing the mass of that object to a standard mass, the terms *weight* and *mass* are sometimes used interchangeably, although this is incorrect.

1.4 Uncertainty in Measurement

The number associated with a measurement is obtained using some measuring device. For example, consider the measurement of the volume of a liquid using a buret (shown in Fig. 1.9 with the scale greatly magnified). Notice that the meniscus of the liquid occurs at about 20.15 milliliters. This means that about 20.15 mL of liquid has been delivered from the buret (if the initial position of the liquid meniscus was 0.00 mL). Note that we must estimate the last number of the volume reading by interpolating between the 0.1-mL marks. Since the last number is estimated, its value may be different if another person makes the same measurement. If five different people read the same volume, the results might be as follows:

Person	Results of Measurement
1	20.15 mL
2	20.14 mL
3	20.16 mL
4	20.17 mL
5	20.16 mL

These results show that the first three numbers (20.1) remain the same regardless of who makes the measurement; these are called *certain* digits. However, the digit to the right of the 1 must be estimated and therefore varies; it is called an *uncertain* digit. We customarily report a measurement by recording all the certain digits plus the *first* uncertain digit. In our example it would not make any sense to try to record the volume of thousandths of a milliliter because the value for hundredths of a milliliter must be estimated when using the buret.

A measurement always has some degree of uncertainty.

It is very important to realize that a *measurement always has some degree of* ***uncertainty.*** The uncertainty of a measurement depends on the precision of the measuring device. For example, using a bathroom scale, you might estimate the mass of a grapefruit to be approximately 1.5 pounds. Weighing the same grapefruit on a highly precise balance might produce a result of 1.476 pounds. In the first case, the uncertainty occurs in the tenths of a pound place; in the second case, the uncertainty occurs in the thousandths of a pound place. Suppose we weigh two similar grapefruits on the two devices and obtain the following results:

	Bathroom Scale	Balance
Grapefruit 1	1.5 lb	1.476 lb
Grapefruit 2	1.5 lb	1.518 lb

Uncertainty in measurement is discussed in more detail in Appendix 1.5.

Do the two grapefruits have the same mass? The answer depends on which set of results you consider. Thus a conclusion based on a series of measurements depends on the certainty of those measurements. For this reason, it is important to indicate the uncertainty in any measurement. This is done by always recording the certain digits and the first uncertain digit (the estimated number). These numbers are called the **significant figures** of a measurement.

The convention of significant figures automatically indicates something about the uncertainty in a measurement. The uncertainty in the last number (the estimated number) is usually assumed to be ±1 unless otherwise indicated. For example, the measurement 1.86 kilograms can be taken to mean 1.86 ± 0.01 kilograms.

Sample Exercise 1.1

Uncertainty in Measurement

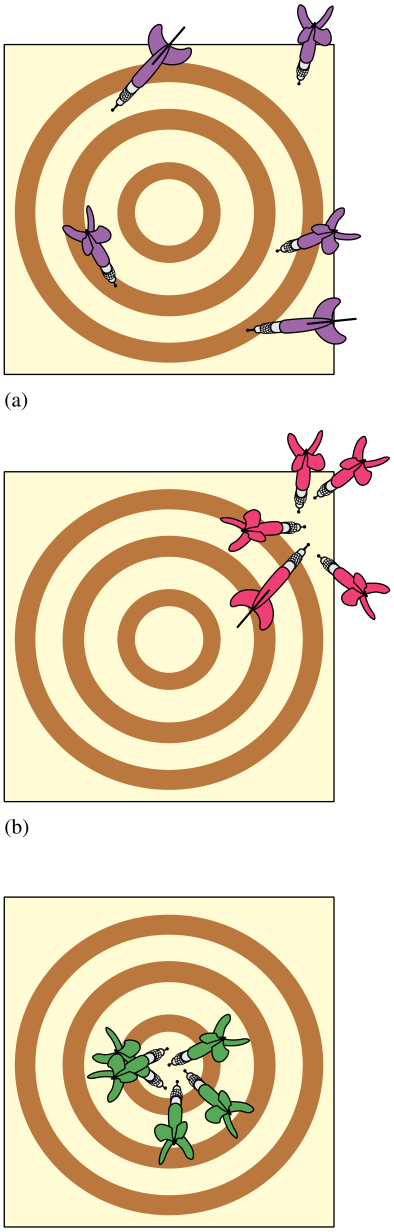
In analyzing a sample of polluted water, a chemist measured out a 25.00-mL water sample with a pipet (see Fig. 1.7). At another point in the analysis, the chemist used a graduated cylinder (see Fig. 1.7) to measure 25 mL of a solution. What is the difference between the measurements 25.00 mL and 25 mL?

Solution

Even though the two volume measurements appear to be equal, they really convey different information. The quantity 25 mL means that the volume is between 24 mL and 26 mL, whereas the quantity 25.00 mL means that the volume is between 24.99 mL and 25.01 mL. The pipet measures volume with much greater precision than does the graduated cylinder.

See Question 1.8.

When making a measurement, it is important to record the results to the appropriate number of significant figures. For example, if a certain buret can be read to ±0.01 mL, you should record a reading of twenty-five milliliters as 25.00 mL, not 25 mL. This way at some later time when you are using your results to do calculations, the uncertainty in the measurement will be known to you.



(a) (b) (c)

FIGURE 1.10
The results of several dart throws show the difference between precise and accurate. (a) Neither accurate nor precise (large random errors). (b) Precise but not accurate (small random errors, large systematic error). (c) Bull's-eye! Both precise and accurate (small random errors, no systematic error).

Precision and Accuracy

Two terms often used to describe the reliability of measurements are *precision* and *accuracy*. Although these words are frequently used interchangeably in everyday life, they have different meanings in the scientific context. **Accuracy** refers to the agreement of a particular value with the true value. **Precision** refers to the degree of agreement among several measurements of the same quantity. Precision reflects the *reproducibility* of a given type of measurement. The difference between these terms is illustrated by the results of three different dart throws shown in Fig. 1.10.

Two different types of errors are illustrated in Fig. 1.10. A **random error** (also called an *indeterminate error*) means that a measurement has an equal probability of being high or low. This type of error occurs in estimating the value of the last digit of a measurement. The second type of error is called **systematic error** (or *determinate error*). This type of error occurs in the same direction each time; it is either always high or always low. Figure 1.10(a) indicates large random errors (poor technique). Figure 1.10(b) indicates small random errors but a large systematic error, and Figure 1.10(c) indicates small random errors and no systematic error.

In quantitative work, precision is often used as an indication of accuracy; we assume that the *average* of a series of precise measurements (which should "average out" the random errors because of their equal probability of being high or low) is accurate, or close to the "true" value. However, this assumption is valid only if systematic errors are absent. Suppose we weigh a piece of brass five times on a very precise balance and obtain the following results:

Weighing	Result
1	2.486 g
2	2.487 g
3	2.485 g
4	2.484 g
5	2.488 g

Normally, we would assume that the true mass of the piece of brass is very close to 2.486 grams, which is the average of the five results:

$$\frac{2.486\text{ g} + 2.487\text{ g} + 2.485\text{ g} + 2.484\text{ g} + 2.488\text{ g}}{5} = 2.486\text{ g}$$

However, if the balance has a defect causing it to give a result that is consistently 1.000 gram too high (a systematic error of $+1.000$ gram), then the measured value of 2.486 grams would be seriously in error. The point here is that high precision among several measurements is an indication of accuracy *only* if systematic errors are absent.

Sample Exercise 1.2 **Precision and Accuracy**

To check the accuracy of a graduated cylinder, a student filled the cylinder to the 25-mL mark using water delivered from a buret (see Fig. 1.7) and then read the volume delivered. Following are the results of five trials:

Trial	Volume Shown by Graduated Cylinder	Volume Shown by the Buret
1	25 mL	26.54 mL
2	25 mL	26.51 mL
3	25 mL	26.60 mL
4	25 mL	26.49 mL
5	25 mL	26.57 mL
Average	25 mL	26.54 mL

Is the graduated cylinder accurate?

Solution

Precision is an indication of accuracy only if there are no systematic errors.

The results of the trials show very good precision (for a graduated cylinder). The student has good technique. However, note that the average value measured using the buret is significantly different from 25 mL. Thus this graduated cylinder is not very accurate. It produces a systematic error (in this case, the indicated result is low for each measurement).

See Question 1.11.

1.5 Significant Figures and Calculations

Calculating the final result for an experiment usually involves adding, subtracting, multiplying, or dividing the results of various types of measurements. Since it is very important that the uncertainty in the final result is known correctly, we have developed rules for counting the significant figures in each number and for determining the correct number of significant figures in the final result.

Rules for Counting Significant Figures

1. *Nonzero integers.* Nonzero integers always count as significant figures.

2. *Zeros*. There are three classes of zeros:
   a. *Leading zeros* are zeros that *precede* all the nonzero digits. These do not count as significant figures. In the number 0.0025, the three zeros simply indicate the position of the decimal point. This number has only two significant figures.
   b. *Captive zeros* are zeros *between* nonzero digits. These always count as significant figures. The number 1.008 has four significant figures.
   c. *Trailing zeros* are zeros at the *right end* of the number. They are significant only if the number contains a decimal point. The number 100 has only one significant figure, whereas the number 1.00×10^2 has three significant figures. The number one hundred written as 100. also has three significant figures.

3. *Exact numbers*. Many times calculations involve numbers that were not obtained using measuring devices but were determined by counting: 10 experiments, 3 apples, 8 molecules. Such numbers are called *exact numbers*. They can be assumed to have an infinite number of significant figures. Other examples of exact numbers are the 2 in $2\pi r$ (the circumference of a circle) and the 4 and the 3 in $\frac{4}{3}\pi r^3$ (the volume of a sphere). Exact numbers also can arise from definitions. For example, one inch is defined as *exactly* 2.54 centimeters. Thus, in the statement 1 in = 2.54 cm, neither the 2.54 nor the 1 limits the number of significant figures when used in a calculation.

Leading zeros are never significant figures.

Captive zeros are always significant figures.

Trailing zeros are sometimes significant figures.

Exact numbers never limit the number of significant figures in a calculation.

Exponential notation is reviewed in Appendix 1.1.

Note that the number 1.00×10^2 above is written in **exponential notation.** This type of notation has at least two advantages: the number of significant figures can be easily indicated, and fewer zeros are needed to write a very large or very small number. For example, the number 0.000060 is much more conveniently represented as 6.0×10^{25}. (The number has two significant figures.)

Sample Exercise 1.3 Significant Figures

Give the number of significant figures for each of the following results.

a. A student's extraction procedure on tea yields 0.0105 g of caffeine.
b. A chemist records a mass of 0.050080 g in an analysis.
c. In an experiment a span of time is determined to be 8.050×10^{-3} s.

Solution

a. The number contains three significant figures. The zeros to the left of the 1 are leading zeros and are not significant, but the remaining zero (a captive zero) is significant.
b. The number contains five significant figures. The leading zeros (to the left of the 5) are not significant. The captive zeros between the 5 and the 8 are significant, and the trailing zero to the right of the 8 is significant because the number contains a decimal point.
c. This number has four significant figures. Both zeros are significant.

See Exercises 1.25 through 1.28.

To this point we have learned to count the significant figures in a given number. Next, we must consider how uncertainty accumulates as calculations are carried out. The detailed analysis of the accumulation of uncertainties depends on the type of calculation involved and can be complex. However, in this textbook we will employ the following simple rules that have been developed for determining the appropriate number of significant figures in the result of a calculation.

Rules for Significant Figures in Mathematical Operations*

1. *For multiplication or division,* the number of significant figures in the result is the same as the number in the least precise measurement used in the calculation. For example, consider the calculation

$$4.56 \times \underset{\substack{\uparrow \\ \text{Limiting term has} \\ \text{two significant} \\ \text{figures}}}{1.4} = 6.38 \xrightarrow{\text{Corrected}} \underset{\substack{\uparrow \\ \text{Two significant} \\ \text{figures}}}{6.4}$$

The product should have only two significant figures, since 1.4 has two significant figures.

2. *For addition or subtraction,* the result has the same number of decimal places as the least precise measurement used in the calculation. For example, consider the sum

$$\begin{array}{rl} 12.11 & \\ 18.0 & \leftarrow \text{Limiting term has one decimal place} \\ \underline{1.013} & \\ 31.123 & \xrightarrow{\text{Corrected}} \underset{\substack{\uparrow \\ \text{One decimal place}}}{31.1} \end{array}$$

The correct result is 31.1, since 18.0 has only one decimal place.

*Although these simple rules work well for most cases, they can give misleading results in certain cases. For more information, see L. M. Schwartz, "Propagation of Significant Figures," *J. Chem. Ed.* **62** (1985): 693; and H. Bradford Thompson, "Is 8°C equal to 50°F?" *J. Chem. Ed.* **68** (1991): 400.

Note that for multiplication and division, significant figures are counted. For addition and subtraction, the decimal places are counted.

In most calculations you will need to round numbers to obtain the correct number of significant figures. The following rules should be applied when rounding.

Rules for Rounding

1. In a series of calculations, carry the extra digits through to the final result, *then* round.
2. If the digit to be removed
   a. is less than 5, the preceding digit stays the same. For example, 1.33 rounds to 1.3.
   b. is equal to or greater than 5, the preceding digit is increased by 1. For example, 1.36 rounds to 1.4.

Rule 2 is consistent with the operation of electronic calculators.

Although rounding is generally straightforward, one point requires special emphasis. As an illustration, suppose that the number 4.348 needs to be rounded to two significant figures. In doing this, we look *only* at the *first number* to the right of the 3:

$$4.3\underset{\uparrow}{4}8$$

Look at this number to round to two significant figures.

Do not round sequentially. The number 6.8347 rounded to three significant figures is 6.83, not 6.84.

The number is rounded to 4.3 because 4 is less than 5. It is incorrect to round sequentially. For example, do *not* round the 4 to 5 to give 4.35 and then round the 3 to 4 to give 4.4.

When rounding, *use only the first number to the right of the last significant figure.*

It is important to note that Rule 1 above usually will not be followed in the Sample Exercises in this text because we want to show the correct number of significant figures in *each step* of a problem. This same practice is followed for the detailed solutions given in the *Solutions Guide.* However, as stated in Rule 1, the best procedure is to carry extra digits throughout a series of calculations and round to the correct number of significant figures only at the end. This is the practice you should follow. The fact that your rounding procedures are different from those used in this text must be taken into account when you check your answer with the one given at the end of the book or in the *Solutions Guide.* Your answer (based on rounding only at the end of a calculation) may differ in the last place from that given here as the “correct” answer because we have rounded after each step. To help you understand the difference between these rounding procedures, we will consider them further in Sample Exercise 1.4.

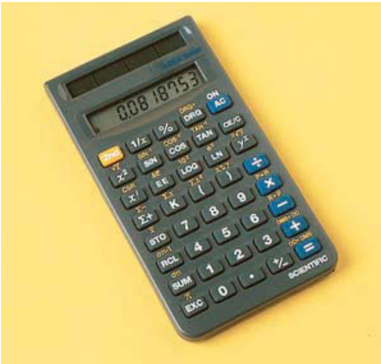
Sample Exercise 1.4 Significant Figures in Mathematical Operations

Carry out the following mathematical operations, and give each result with the correct number of significant figures.

a. $1.05 \times 10^{-3} \div 6.135$
b. $21 - 13.8$
c. As part of a lab assignment to determine the value of the gas constant (R), a student measured the pressure (P), volume (V), and temperature (T) for a sample of gas, where

$$R = \frac{PV}{T}$$

The following values were obtained: $P = 2.560$, $T = 275.15$, and $V = 8.8$. (Gases will be discussed in detail in Chapter 5; we will not be concerned at this time about the units for these quantities.) Calculate R to the correct number of significant figures.



This number must be rounded to two significant figures.

Solution

a. The result is 1.71×10^{-4}, which has three significant figures because the term with the least precision (1.05×10^{-3}) has three significant figures.

b. The result is 7 with no decimal point because the number with the least number of decimal places (21) has none.

c. $$R = \frac{PV}{T} = \frac{(2.560)(8.8)}{275.15}$$

The correct procedure for obtaining the final result can be represented as follows:

$$\frac{(2.560)(8.8)}{275.15} = \frac{22.528}{275.15} = 0.0818753$$
$$= 0.082 = 8.2 \times 10^{-2} = R$$

The final result must be rounded to two significant figures because 8.8 (the least precise measurement) has two significant figures. To show the effects of rounding at intermediate steps, we will carry out the calculation as follows:

$$\frac{(2.560)(8.8)}{275.15} = \frac{22.528}{275.15} = \frac{\overset{\substack{\text{Rounded to two} \\ \text{significant figures} \\ \downarrow}}{23}}{275.15}$$

Now we proceed with the next calculation:

$$\frac{23}{275.15} = 0.0835908$$

Rounded to two significant figures, this result is

$$0.084 = 8.4 \times 10^{-2}$$

Note that intermediate rounding gives a significantly different result than was obtained by rounding only at the end. Again, we must reemphasize that in *your* calculations you should round *only at the end.* However, because rounding is carried out at intermediate steps in this text (to always show the correct number of significant figures), the final answer given in the text may differ slightly from the one you obtain (rounding only at the end).

See Exercises 1.31 through 1.34.

There is a useful lesson to be learned from part c of Sample Exercise 1.4. The student measured the pressure and temperature to greater precision than the volume. A more precise value of R (one with more significant figures) could have been obtained if a more precise measurement of V had been made. As it is, the efforts expended to measure P and T very precisely were wasted. Remember that a series of measurements to obtain some final result should all be done to about the same precision.

TABLE 1.4 English–Metric Equivalents

Length	1 m = 1.094 yd
	2.54 cm = 1 in
Mass	1 kg = 2.205 lb
	453.6 g = 1 lb
Volume	1 L = 1.06 qt
	1 ft^3 = 28.32 L

1.6 Dimensional Analysis

It is often necessary to convert a given result from one system of units to another. The best way to do this is by a method called the **unit factor method,** or more commonly **dimensional analysis.** To illustrate the use of this method, we will consider several unit conversions. Some equivalents in the English and metric systems are listed in Table 1.4. A more complete list of conversion factors given to more significant figures appears in Appendix 6.

Consider a pin measuring 2.85 centimeters in length. What is its length in inches? To accomplish this conversion, we must use the equivalence statement

$$2.54 \text{ cm} = 1 \text{ in}$$

If we divide both sides of this equation by 2.54 centimeters, we get

$$1 = \frac{1 \text{ in}}{2.54 \text{ cm}}$$

This expression is called a *unit factor.* Since 1 inch and 2.54 centimeters are exactly equivalent, multiplying any expression by this unit factor will not change its *value.*

The pin has a length of 2.85 centimeters. Multiplying this length by the appropriate unit factor gives

$$2.85 \cancel{\text{cm}} \times \frac{1 \text{ in}}{2.54 \cancel{\text{cm}}} = \frac{2.85}{2.54} \text{ in} = 1.12 \text{ in}$$

Note that the centimeter units cancel to give inches for the result. This is exactly what we wanted to accomplish. Note also that the result has three significant figures, as required by the number 2.85. Recall that the 1 and 2.54 in the conversion factor are exact numbers by definition.

Sample Exercise 1.5

Unit Conversions I

A pencil is 7.00 in long. What is its length in centimeters?

Solution

In this case we want to convert from inches to centimeters. Therefore, we must use the reciprocal of the unit factor used above to do the opposite conversion:

$$7.00 \cancel{\text{in}} \times \frac{2.54 \text{ cm}}{1 \cancel{\text{in}}} = (7.00)(2.54) \text{ cm} = 17.8 \text{ cm}$$

Here the inch units cancel, leaving centimeters, as requested.

See Exercises 1.37 and 1.38.

Note that two unit factors can be derived from each equivalence statement. For example, from the equivalence statement 2.54 cm = 1 in, the two unit factors are

$$\frac{2.54 \text{ cm}}{1 \text{ in}} \quad \text{and} \quad \frac{1 \text{ in}}{2.54 \text{ cm}}$$

Consider the direction of the required change to select the correct unit factor.

How do you choose which one to use in a given situation? Simply look at the *direction* of the required change. To change from inches to centimeters, the inches must cancel. Thus the factor 2.54 cm/1 in is used. To change from centimeters to inches, centimeters must cancel, and the factor 1 in/2.54 cm is appropriate.

Converting from One Unit to Another

- To convert from one unit to another, use the equivalence statement that relates the two units.
- Derive the appropriate unit factor by looking at the direction of the required change (to cancel the unwanted units).
- Multiply the quantity to be converted by the unit factor to give the quantity with the desired units.

Sample Exercise 1.6 Unit Conversions II

You want to order a bicycle with a 25.5-in frame, but the sizes in the catalog are given only in centimeters. What size should you order?

Solution

You need to go from inches to centimeters, so 2.54 cm = 1 in is appropriate:

$$25.5\ \cancel{\text{in}} \times \frac{2.54\ \text{cm}}{1\ \cancel{\text{in}}} = 64.8\ \text{cm}$$

See Exercises 1.37 and 1.38.

To ensure that the conversion procedure is clear, a multistep problem is considered in Sample Exercise 1.7.

Sample Exercise 1.7 Unit Conversions III

A student has entered a 10.0-km run. How long is the run in miles?

Solution

This conversion can be accomplished in several different ways. Since we have the equivalence statement 1 m = 1.094 yd, we will proceed by a path that uses this fact. Before we start any calculations, let us consider our strategy. We have kilometers, which we want to change to miles. We can do this by the following route:

$$\text{kilometers} \longrightarrow \text{meters} \longrightarrow \text{yards} \longrightarrow \text{miles}$$

To proceed in this way, we need the following equivalence statements:

$$1\ \text{km} = 1000\ \text{m}$$
$$1\ \text{m} = 1.094\ \text{yd}$$
$$1760\ \text{yd} = 1\ \text{mi}$$

To make sure the process is clear, we will proceed step by step:

Kilometers to Meters:

$$10.0\ \cancel{\text{km}} \times \frac{1000\ \text{m}}{1\ \cancel{\text{km}}} = 1.00 \times 10^4\ \text{m}$$

Meters to Yards:

$$1.00 \times 10^4\ \cancel{\text{m}} \times \frac{1.094\ \text{yd}}{1\ \cancel{\text{m}}} = 1.094 \times 10^4\ \text{yd}$$

Note that we should have only three significant figures in the result. Since this is an intermediate result, however, we will carry the extra digit. Remember, round off only the final result.

Yards to Miles:

Normally we round to the correct number of significant figures after each step. However, you should round only at the end.

$$1.094 \times 10^4\ \cancel{\text{yd}} \times \frac{1\ \text{mi}}{1760\ \cancel{\text{yd}}} = 6.216\ \text{mi}$$

Note in this case that 1 mi equals exactly 1760 yd *by designation*. Thus 1760 is an exact number.

Since the distance was originally given as 10.0 km, the result can have only three significant figures and should be rounded to 6.22 mi. Thus

$$10.0\ \text{km} = 6.22\ \text{mi}$$

Alternatively, we can combine the steps:

$$10.0\ \cancel{\text{km}} \times \frac{1000\ \cancel{\text{m}}}{1\ \cancel{\text{km}}} \times \frac{1.094\ \cancel{\text{yd}}}{1\ \cancel{\text{m}}} \times \frac{1\ \text{mi}}{1760\ \cancel{\text{yd}}} = 6.22\ \text{mi}$$

See Exercises 1.37 and 1.38.

In using dimensional analysis, your verification that everything has been done correctly is that you end up with the correct units. *In doing chemistry problems, you should always include the units for the quantities used*. Always check to see that the units cancel to give the correct units for the final result. This provides a very valuable check, especially for complicated problems.

Study the procedures for unit conversions in the following Sample Exercises.

Sample Exercise 1.8

Unit Conversion IV

The speed limit on many highways in the United States is 55 mi/h. What number would be posted in kilometers per hour?

Solution

Result obtained by rounding only at the end of the calculation

$$\frac{55\ \cancel{\text{mi}}}{\text{h}} \times \frac{1760\ \cancel{\text{yd}}}{1\ \cancel{\text{mi}}} \times \frac{1\ \cancel{\text{m}}}{1.094\ \cancel{\text{yd}}} \times \frac{1\ \text{km}}{1000\ \cancel{\text{m}}} = 88\ \text{km/h}$$

Note that all units cancel except the desired kilometers per hour.

See Exercises 1.43 through 1.45.

Sample Exercise 1.9

Unit Conversions V

A Japanese car is advertised as having a gas mileage of 15 km/L. Convert this rating to miles per gallon.

Solution

Result obtained by rounding only at the end of the calculation

$$\frac{15\ \cancel{\text{km}}}{\cancel{\text{L}}} \times \frac{1000\ \cancel{\text{m}}}{1\ \cancel{\text{km}}} \times \frac{1.094\ \cancel{\text{yd}}}{1\ \cancel{\text{m}}} \times \frac{1\ \text{mi}}{1760\ \cancel{\text{yd}}} \times \frac{1\ \cancel{\text{L}}}{1.06\ \cancel{\text{qt}}} \times \frac{4\ \cancel{\text{qt}}}{1\ \text{gal}} = 35\ \text{mi/gal}$$

See Exercise 1.46.

1.7 Temperature

Three systems for measuring temperature are widely used: the Celsius scale, the Kelvin scale, and the Fahrenheit scale. The first two temperature systems are used in the physical sciences, and the third is used in many of the engineering sciences. Our purpose here is to define the three temperature scales and show how conversions from one scale to another can be performed. Although these conversions can be carried out routinely on most calculators, we will consider the process in some detail here to illustrate methods of problem solving.

The three temperature scales are defined and compared in Fig. 1.11. Note that the size of the temperature unit (the *degree*) is the same for the Kelvin and Celsius scales.

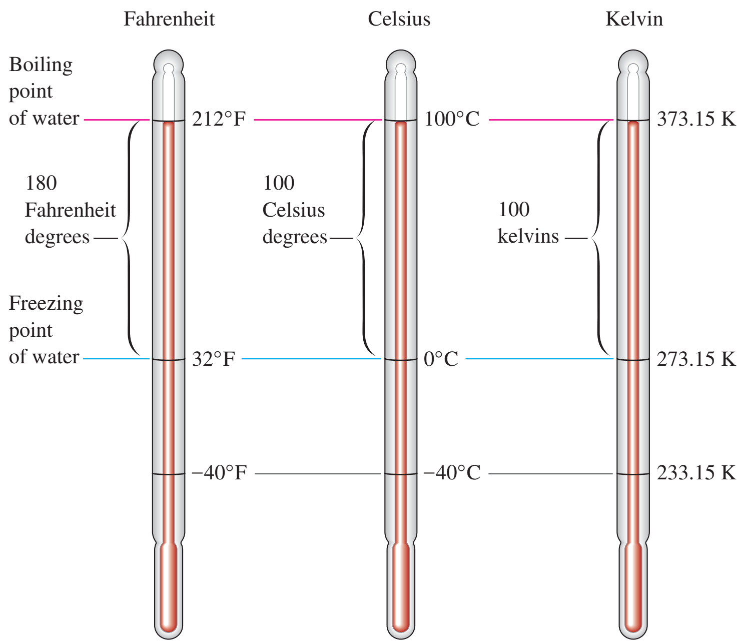


FIGURE 1.11
The three major temperature scales.

The fundamental difference between these two temperature scales is in their zero points. Conversion between these two scales simply requires an adjustment for the different zero points.

$T_K = T_C + 273.15$

$$\text{Temperature (Kelvin)} = \text{temperature (Celsius)} + 273.15$$

or

$T_C = T_K - 273.15$

$$\text{Temperature (Celsius)} = \text{temperature (Kelvin)} - 273.15$$

For example, to convert 300.00 K to the Celsius scale, we do the following calculation:

$$300.00 - 273.15 = 26.85°\text{C}$$

Note that in expressing temperature in Celsius units, the designation °C is used. The degree symbol is not used when writing temperature in terms of the Kelvin scale. The unit of temperature on this scale is called a *kelvin* and is symbolized by the letter K.

Converting between the Fahrenheit and Celsius scales is somewhat more complicated because both the degree sizes and the zero points are different. Thus we need to consider two adjustments: one for degree size and one for the zero point. First, we must account for the difference in degree size. This can be done by reconsidering Fig. 1.11. Notice that since 212°F = 100°C and 32°F = 0°C,

$$212 - 32 = 180 \text{ Fahrenheit degrees} = 100 - 0 = 100 \text{ Celsius degrees}$$

Thus 180° on the Fahrenheit scale is equivalent to 100° on the Celsius scale, and the unit factor is

$$\frac{180°\text{F}}{100°\text{C}} \quad \text{or} \quad \frac{9°\text{F}}{5°\text{C}}$$

or the reciprocal, depending on the direction in which we need to go.

Next, we must consider the different zero points. Since 32°F = 0°C, we obtain the corresponding Celsius temperature by first subtracting 32 from the Fahrenheit temperature to account for the different zero points. Then the unit factor is applied to adjust for the difference in the degree size. This process is summarized by the equation

$$(T_F - 32°F)\frac{5°C}{9°F} = T_C \tag{1.1}$$

where T_F and T_C represent a given temperature on the Fahrenheit and Celsius scales, respectively. In the opposite conversion, we first correct for degree size and then correct for the different zero point. This process can be summarized in the following general equation:

$$T_F = T_C \times \frac{9°F}{5°C} + 32°F \tag{1.2}$$

Equations (1.1) and (1.2) are really the same equation in different forms. See if you can obtain Equation (1.2) by starting with Equation (1.1) and rearranging.

Understand the process of converting from one temperature scale to another; do not simply memorize the equations.

At this point it is worthwhile to weigh the two alternatives for learning to do temperature conversions: You can simply memorize the equations, or you can take the time to learn the differences between the temperature scales and to understand the processes involved in converting from one scale to another. The latter approach may take a little more effort, but the understanding you gain will stick with you much longer than the memorized formulas. This choice also will apply to many of the other chemical concepts. Try to think things through!

Sample Exercise 1.10 Temperature Conversions I

Normal body temperature is 98.6°F. Convert this temperature to the Celsius and Kelvin scales.

Solution

Rather than simply using the formulas to solve this problem, we will proceed by thinking it through. The situation is diagramed in Fig. 1.12. First, we want to convert 98.6°F to the Celsius scale. The number of Fahrenheit degrees between 32.0°F and 98.6°F is 66.6°F. We must convert this difference to Celsius degrees:

$$66.6°F \times \frac{5°C}{9°F} = 37.0°C$$

A physician taking the temperature of a patient.

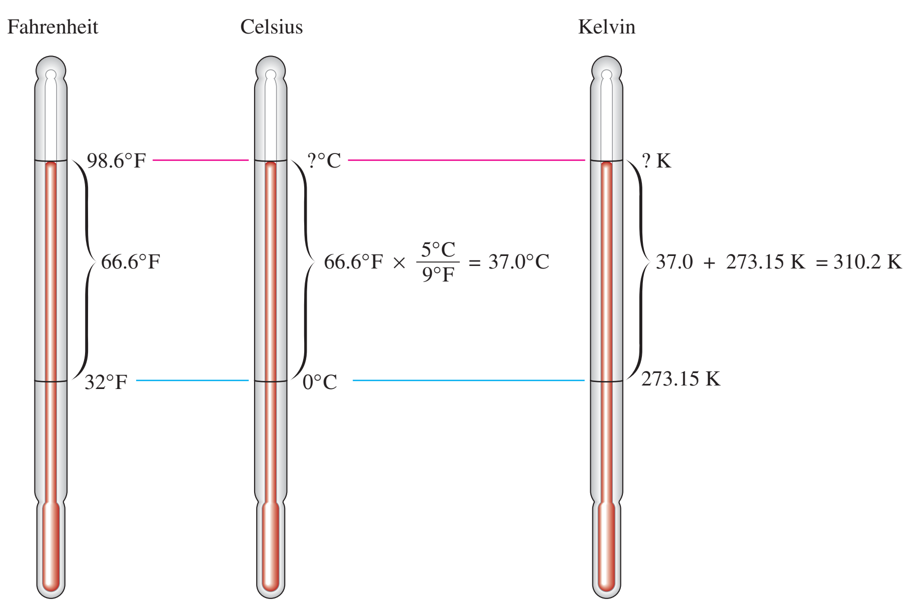


FIGURE 1.12
Normal body temperature on the Fahrenheit, Celsius, and Kelvin scales.

CHEMICAL IMPACT

Faux Snow

Skiing is challenging and fun, but it is also big business. Both skiers and ski operators want the season to last as long as possible. The major factor in maximizing the length of the ski season and in salvaging dry periods during the winter is the ability to "make snow." Machine-made snow is now a required part of maintaining ideal conditions at major ski areas such as Aspen, Breckenridge, and Taos.

Snow is relatively easy to make if the air is cold enough. To manufacture snow, water is cooled to just above 0°C and then pumped at high pressure through a "gun" that produces a fine mist of water droplets that freeze before dropping to the ground. As might be expected, atmospheric conditions are critical when making snow. With an air temperature of −8°C (18°F) or less, untreated water can be used in the snow guns. However, the ideal type of snow for skiing is "powder"—fluffy snow made up of small, individual crystals. To achieve powdery snow requires sufficient nucleation sites—that is, sites where crystal growth is initiated. This condition can be achieved by "doping" the water with ions such as calcium or magnesium or with fine particles of clay. Also, when the air temperature is between 0°C and −8°C, materials such as silver iodide, detergents, and organic materials may be added to the water to seed the snow.

A discovery at the University of Wisconsin in the 1970s led to the additive most commonly used for snow making. The Wisconsin scientists found that a bacterium (*Pseudomanas syringae*) commonly found in nature makes a protein that acts as a very effective nucleation site for ice formation. In fact, this discovery helped to explain why ice forms at 0°C on the blossoms of fruit trees instead of the water supercooling below 0°C, as pure water does when the temperature is lowered slowly below the freezing point. To help protect fruit blossoms from freeze damage, this bacterium has been genetically modified to remove the ice nucleation protein. As a result, fruit blossoms can survive

Thus 98.6°F corresponds to 37.0°C.

Now we can convert to the Kelvin scale:

$$T_K = T_C + 273.15 = 37.0 + 273.15 = 310.2 \text{ K}$$

Note that the final answer has only one decimal place (37.0 is limiting).

See Exercises 1.49, 1.51, and 1.52.

Sample Exercise 1.11

Temperature Conversions II

One interesting feature of the Celsius and Fahrenheit scales is that −40°C and −40°F represent the same temperature, as shown in Fig. 1.11. Verify that this is true.

Solution

The difference between 32°F and −40°F is 72°F. The difference between 0°C and −40°C is 40°C. The ratio of these is

$$\frac{72°\text{F}}{40°\text{C}} = \frac{8 \times 9°\text{F}}{8 \times 5°\text{C}} = \frac{9°\text{F}}{5°\text{C}}$$

as required. Thus −40°C is equivalent to −40°F.

See Challenge Problem 1.86.

Since, as shown in Sample Exercise 1.11, −40° on both the Fahrenheit and Celsius scales represents the same temperature, this point can be used as a reference point (like 0°C and 32°F) for a relationship between the two scales:

$$\frac{\text{Number of Fahrenheit degrees}}{\text{Number of Celsius degrees}} = \frac{T_F - (-40)}{T_C - (-40)} = \frac{9°\text{F}}{5°\text{C}}$$

intact even if the temperature briefly falls below 0°C. (See the Chemical Impact on Organisms and Ice Formation on page 516.) For snow-making purposes, this protein forms the basis for Snowmax (prepared and sold by York Snow of Victor, New York), which is the most popular additive for snow making.

Obviously, snow cannot be made in the summer, so what is a skiing fanatic to do during the warm months? The answer is "dryslope" skiing. Although materials for dryslopes can be manufactured in a variety of ways, polymers are most commonly used for this application. One company that makes a multilayer polymer for artificial ski slopes is Briton Engineering Developments (Yorkshire, England), the producer of Snowflex. Snowflex consists of a slippery polymer fiber placed on top of a shock-absorbing base and lubricated by misting water through holes in its surface. Of course, this virtual skiing is not much like the real thing but it does provide some relief for summer ski withdrawal.

As artificial and synthetic snow amply demonstrate, chemistry makes life more fun.

A freestyle ski area at Sheffield Ski Village, in England, uses Snowflex "virtual snow" for year-round fun.

or
$$\frac{T_F + 40}{T_C + 40} = \frac{9°F}{5°C} \tag{1.3}$$

where T_F and T_C represent the same temperature (but not the same number). This equation can be used to convert Fahrenheit temperatures to Celsius, and vice versa, and may be easier to remember than Equations (1.1) and (1.2).

Sample Exercise 1.12

Temperature Conversions III

Liquid nitrogen, which is often used as a coolant for low-temperature experiments, has a boiling point of 77 K. What is this temperature on the Fahrenheit scale?

Liquid nitrogen is so cold that water condenses out of the surrounding air, forming a cloud as the nitrogen is poured.

Solution

We will first convert 77 K to the Celsius scale:

$$T_C = T_K - 273.15 = 77 - 273.15 = -196°C$$

To convert to the Fahrenheit scale, we will use Equation (1.3):

$$\frac{T_F + 40}{T_C + 40} = \frac{9°F}{5°C}$$

$$\frac{T_F + 40}{-196°C + 40} = \frac{T_F + 40}{-156°C} = \frac{9°F}{5°C}$$

$$T_F + 40 = \frac{9°F}{5°C}(-156°C) = -281°F$$

$$T_F = -281°F - 40 = -321°F$$

See Exercises 1.49, 1.51, and 1.52.

1.8 Density

A property of matter that is often used by chemists as an "identification tag" for a substance is **density,** the mass of substance per unit volume of the substance:

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

The density of a liquid can be determined easily by weighing an accurately known volume of liquid. This procedure is illustrated in Sample Exercise 1.13.

Sample Exercise 1.13 Determining Density

A chemist, trying to identify the main component of a compact disc cleaning fluid, finds that 25.00 cm^3 of the substance has a mass of 19.625 g at 20°C. The following are the names and densities of the compounds that might be the main component:

Compound	Density in g/cm^3 at 20°C
Chloroform	1.492
Diethyl ether	0.714
Ethanol	0.789
Isopropyl alcohol	0.785
Toluene	0.867

Which of these compounds is the most likely to be the main component of the compact disc cleaner?

Solution

There are two ways of indicating units that occur in the denominator. For example, we can write g/cm^3 or $g\ cm^{-3}$. Although we will use the former system here, the other system is widely used.

To identify the unknown substance, we must determine its density. This can be done by using the definition of density:

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{19.625\ \text{g}}{25.00\ \text{cm}^3} = 0.7850\ \text{g/cm}^3$$

This density corresponds exactly to that of isopropyl alcohol, which is therefore the most likely main component of the cleaner. However, note that the density of ethanol is also very close. To be sure that the compound is isopropyl alcohol, we should run several more density experiments. (In the modern laboratory, many other types of tests could be done to distinguish between these two liquids.)

See Exercises 1.55 and 1.56.

Besides being a tool for the identification of substances, density has many other uses. For example, the liquid in your car's lead storage battery (a solution of sulfuric acid) changes density because the sulfuric acid is consumed as the battery discharges. In a fully charged battery, the density of the solution is about 1.30 g/cm^3. If the density falls below 1.20 g/cm^3, the battery will have to be recharged. Density measurement is also used to determine the amount of antifreeze, and thus the level of protection against freezing, in the cooling system of a car.

The densities of various common substances are given in Table 1.5.

TABLE 1.5 Densities of Various Common Substances* at 20°C

Substance	Physical State	Density (g/cm^3)
Oxygen	Gas	0.00133
Hydrogen	Gas	0.000084
Ethanol	Liquid	0.789
Benzene	Liquid	0.880
Water	Liquid	0.9982
Magnesium	Solid	1.74
Salt (sodium chloride)	Solid	2.16
Aluminum	Solid	2.70
Iron	Solid	7.87
Copper	Solid	8.96
Silver	Solid	10.5
Lead	Solid	11.34
Mercury	Liquid	13.6
Gold	Solid	19.32

*At 1 atmosphere pressure

1.9 Classification of Matter

Before we can hope to understand the changes we see going on around us—the growth of plants, the rusting of steel, the aging of people, rain becoming more acidic—we must find out how matter is organized. **Matter,** best defined as anything occupying space and having mass, is the material of the universe. Matter is complex and has many levels of organization. In this section we introduce basic ideas about the structure of matter and its behavior.

Visualizations:
Structure of a Gas
Structure of a Liquid
Structure of a Solid

We will start by considering the definitions of the fundamental properties of matter. Matter exists in three **states:** solid, liquid, and gas. A *solid* is rigid; it has a fixed volume and shape. A *liquid* has a definite volume but no specific shape; it assumes the shape of its container. A *gas* has no fixed volume or shape; it takes on the shape and volume of its container. In contrast to liquids and solids, which are only slightly compressible, gases are highly compressible; it is relatively easy to decrease the volume of a gas. Molecular-level pictures of the three states of water are given in Fig. 1.13. The different properties of ice, liquid water, and steam are determined by the different arrangements of the molecules in these substances. Table 1.5 gives the states of some common substances at 20°C and 1 atmosphere of pressure.

Visualization: Comparison of a Compound and a Mixture

Most of the matter around us consists of **mixtures** of pure substances. Wood, gasoline, wine, soil, and air are all mixtures. The main characteristic of a mixture is that it has *variable composition.* For example, wood is a mixture of many substances, the proportions of which vary depending on the type of wood and where it grows. Mixtures can be classified as **homogeneous** (having visibly indistinguishable parts) or **heterogeneous** (having visibly distinguishable parts).

A homogeneous mixture is called a **solution.** Air is a solution consisting of a mixture of gases. Wine is a complex liquid solution. Brass is a solid solution of copper and zinc. Sand in water and iced tea with ice cubes are examples of heterogeneous mixtures. Heterogeneous mixtures usually can be separated into two or more homogeneous mixtures or pure substances (for example, the ice cubes can be separated from the tea).

Visualization: Comparison of a Solution and a Mixture

Visualization: Homogeneous Mixtures: Air and Brass)

Mixtures can be separated into pure substances by physical methods. A **pure substance** is one with constant composition. Water is a good illustration of these ideas. As we will discuss in detail later, pure water is composed solely of H_2O molecules,

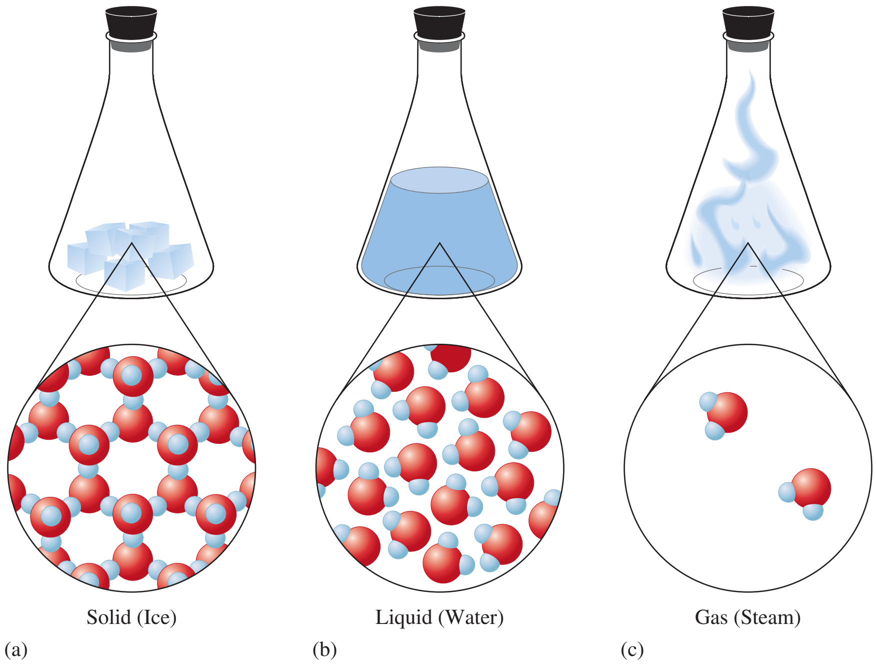


FIGURE 1.13
The three states of water (where red spheres represent oxygen atoms and blue spheres represent hydrogen atoms). (a) Solid: the water molecules are locked into rigid positions and are close together. (b) Liquid: the water molecules are still close together but can move around to some extent. (c) Gas: the water molecules are far apart and move randomly.

but the water found in nature (groundwater or the water in a lake or ocean) is really a mixture. Seawater, for example, contains large amounts of dissolved minerals. Boiling seawater produces steam, which can be condensed to pure water, leaving the minerals behind as solids. The dissolved minerals in seawater also can be separated out by freezing the mixture, since pure water freezes out. The processes of boiling and freezing are **physical changes:** When water freezes or boils, it changes its state but remains water; it is still composed of H_2O molecules. A physical change is a change in the form of a substance, not in its chemical composition. A physical change can be used to separate a mixture into pure compounds, but it will not break compounds into elements.

One of the most important methods for separating the components of a mixture is **distillation,** a process that depends on differences in the volatility (how readily substances become gases) of the components. In simple distillation, a mixture is heated in a device such as that shown in Fig. 1.14. The most volatile component vaporizes at the lowest temperature, and the vapor passes through a cooled tube (a condenser), where it condenses back into its liquid state.

The term *volatile* refers to the ease with which a substance can be changed to its vapor.

The simple, one-stage distillation apparatus shown in Fig. 1.14 works very well when only one component of the mixture is volatile. For example, a mixture of water and sand is easily separated by boiling off the water. Water containing dissolved minerals behaves in much the same way. As the water is boiled off, the minerals remain behind as nonvolatile solids. Simple distillation of seawater using the sun as the heat source is an excellent way to desalinate (remove the minerals from) seawater.

However, when a mixture contains several volatile components, the one-step distillation does not give a pure substance in the receiving flask, and more elaborate methods are required.

Another method of separation is simple **filtration,** which is used when a mixture consists of a solid and a liquid. The mixture is poured onto a mesh, such as filter paper, which passes the liquid and leaves the solid behind.

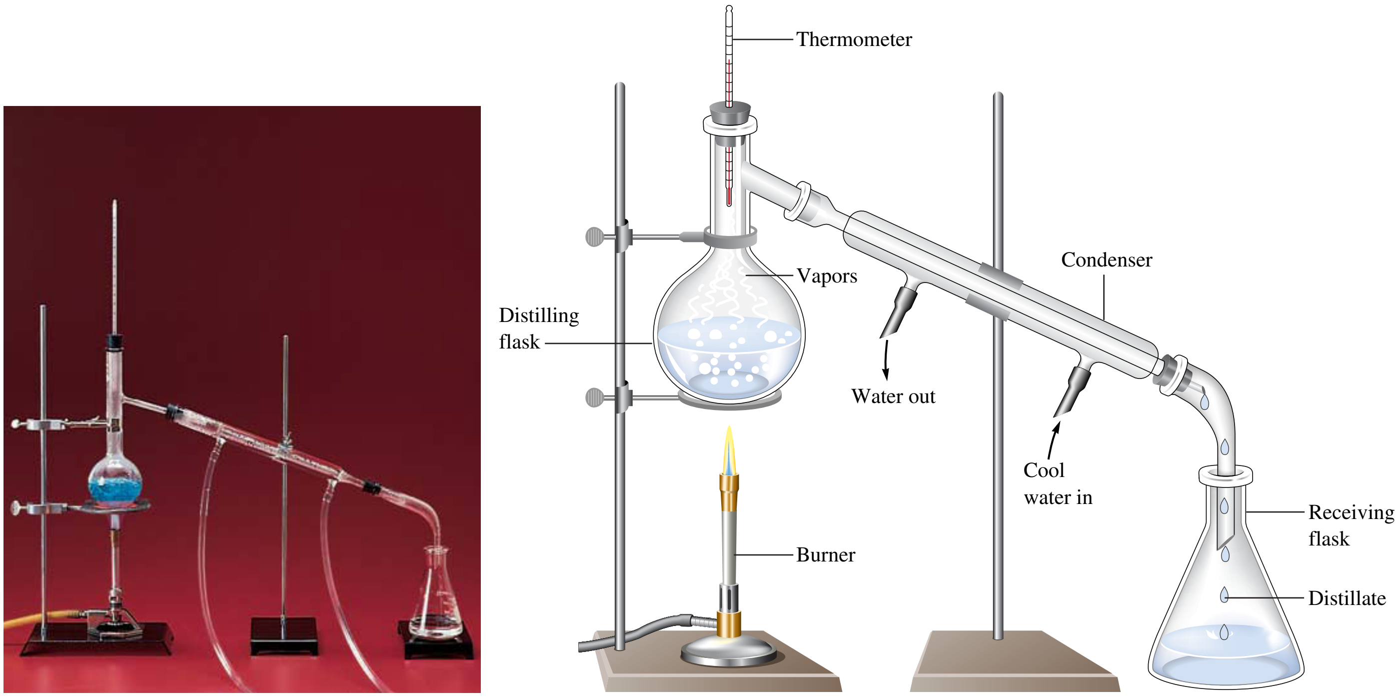


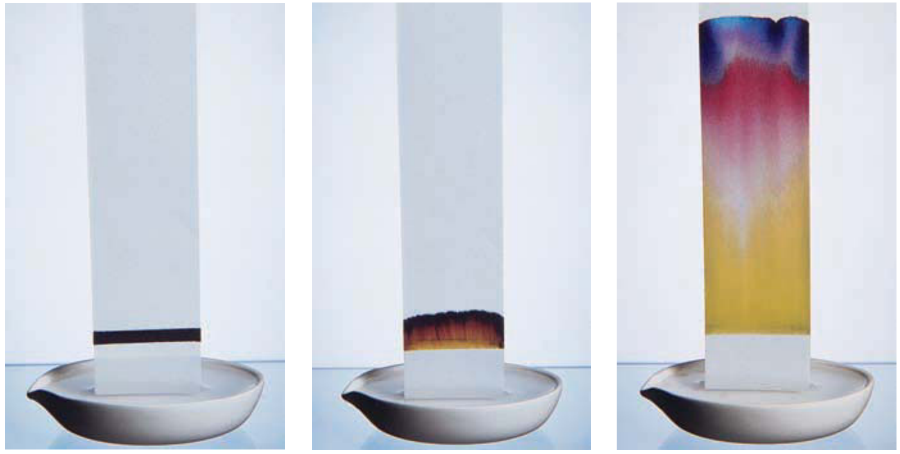
FIGURE 1.14
Simple laboratory distillation apparatus. Cool water circulates through the outer portion of the condenser, causing vapors from the distilling flask to condense into a liquid. The nonvolatile component of the mixture remains in the distilling flask.

A third method of separation is called **chromatography.** Chromatography is the general name applied to a series of methods that employ a system with two *phases* (states) of matter: a mobile phase and a stationary phase. The *stationary phase* is a solid, and the *mobile phase* is either a liquid or a gas. The separation process occurs because the components of the mixture have different affinities for the two phases and thus move through the system at different rates. A component with a high affinity for the mobile phase moves relatively quickly through the chromatographic system, whereas one with a high affinity for the solid phase moves more slowly.

One simple type of chromatography, **paper chromatography,** employs a strip of porous paper, such as filter paper, for the stationary phase. A drop of the mixture to be separated is placed on the paper, which is then dipped into a liquid (the mobile phase) that travels up the paper as though it were a wick (see Fig. 1.15). This method of separating a mixture is often used by biochemists, who study the chemistry of living systems.

It should be noted that when a mixture is separated, the absolute purity of the separated components is an ideal. Because water, for example, inevitably comes into contact with other materials when it is synthesized or separated from a mixture, it is never absolutely pure. With great care, however, substances can be obtained in very nearly pure form.

Pure substances contain compounds (combinations of elements) or free elements. A **compound** is a substance with *constant composition* that can be broken down into elements by chemical processes. An example of a chemical process is the electrolysis of water, in which an electric current is passed through water to break it down into the free elements hydrogen and oxygen. This process produces a chemical change because the water molecules have been broken down. The water is gone, and in its place we have the free elements hydrogen and oxygen. A **chemical change** is one in which a given substance becomes a new substance or substances with different properties and different



(a) (b) (c)

FIGURE 1.15
Paper chromatography of ink. (a) A line of the mixture to be separated is placed at one end of a sheet of porous paper. (b) The paper acts as a wick to draw up the liquid. (c) The component with the weakest attraction for the paper travels faster than the components that cling to the paper.

composition. **Elements** are substances that cannot be decomposed into simpler substances by chemical or physical means.

We have seen that the matter around us has various levels of organization. The most fundamental substances we have discussed so far are elements. As we will see in later chapters, elements also have structure: They are composed of atoms, which in turn are composed of nuclei and electrons. Even the nucleus has structure: It is composed of protons and neutrons. And even these can be broken down further, into elementary particles called *quarks*. However, we need not concern ourselves with such details at this point. Figure 1.16 summarizes our discussion of the organization of matter.

The element mercury (top left) combines with the element iodine (top right) to form the compound mercuric iodide (bottom). This is an example of a chemical change.

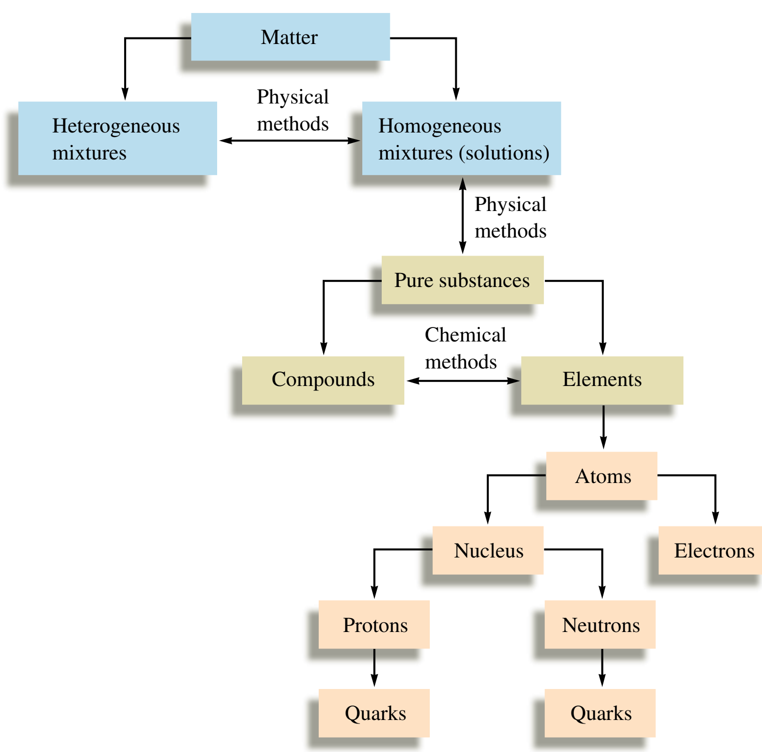


FIGURE 1.16
The organization of matter.

Key Terms

Section 1.2
scientific method
measurement
hypothesis
theory
model
natural law
law of conservation of mass

Section 1.3
SI system
mass
weight

Section 1.4
uncertainty
significant figures
accuracy
precision
random error
systematic error

Section 1.5
exponential notation

Section 1.6
unit factor method
dimensional analysis

Section 1.8
density

Section 1.9
matter
states (of matter)
homogeneous mixture
heterogeneous mixture
solution
pure substance
physical change
distillation
filtration
chromatography
paper chromatography
compound
chemical change
element

For Review

Scientific method
- Make observations
- Formulate hypotheses
- Perform experiments

Models (theories) are explanation of why nature behaves in a particular way.
- They are subject to modification over time and sometimes fail.

Quantitative observations are called measurements.
- Consist of a number and a unit
- Involve some uncertainty
- Uncertainty is indicated by using significant figures
  - Rules to determine significant figures
  - Calculations using significant figures
- Preferred system is SI

Temperature conversions
- $T_K = T_C + 273$
- $T_C = (T_F - 32°F)\left(\frac{5°C}{9°F}\right)$
- $T_F = T_C\left(\frac{5°F}{9°C}\right) + 32°F$

Density
- $\text{Density} = \frac{\text{mass}}{\text{volume}}$

Matter can exist in three states:
- Solid
- Liquid
- Gas

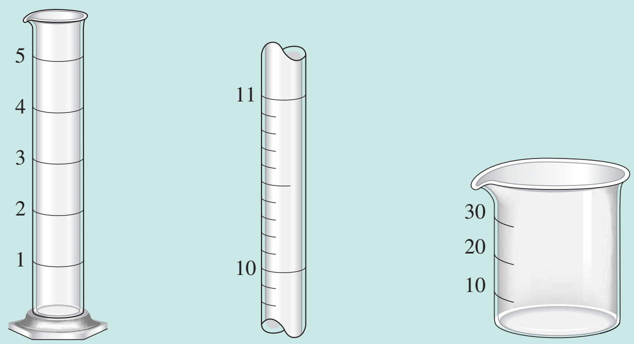
Mixtures can be separated by methods involving only physical changes:
- Distillation
- Filtration
- Chromatography

Compounds can be decomposed to elements only through chemical changes.

REVIEW QUESTIONS

1. Define and explain the differences between the following terms.
   a. law and theory
   b. theory and experiment
   c. qualitative and quantitative
   d. hypothesis and theory
2. Is the scientific method suitable for solving problems only in the sciences? Explain.
3. Which of the following statements (hypotheses) could be tested by quantitative measurement?
   a. Ty Cobb was a better hitter than Pete Rose.
   b. Ivory soap is $99\frac{44}{100}\%$ pure.
   c. Rolaids consumes 47 times its weight in excess stomach acid.

4. For each of the following pieces of glassware, provide a sample measurement and discuss the number of significant figures and uncertainty.



a. b. c.

5. A student performed an analysis of a sample for its calcium content and got the following results:

14.92% 14.91% 14.88% 14.91%

The actual amount of calcium in the sample is 15.70%. What conclusions can you draw about the accuracy and precision of these results?

6. Compare and contrast the multiplication/division significant figure rule to the significant figure rule applied for addition/subtraction mathematical operations.
7. Explain how density can be used as a conversion factor to convert the volume of an object to the mass of the object, and vice versa.
8. On which temperature scale (°F, °C, or K) does 1 degree represent the smallest change in temperature?
9. Distinguish between physical changes and chemical changes.
10. Why is the separation of mixtures into pure or relatively pure substances so important when performing a chemical analysis?

Active Learning Questions

These questions are designed to be used by groups of students in class. The questions allow students to explore their understanding of concepts through discussion and peer teaching. The real value of these questions is the learning that occurs while students talk to each other about chemical concepts.

1. a. There are 365 days per year, 24 hours per day, 12 months per year, and 60 minutes per hour. Use these data to determine how many minutes are in a month.

b. Now use the following data to calculate the number of minutes in a month: 24 hours per day, 60 minutes per hour, 7 days per week, and 4 weeks per month.

c. Why are these answers different? Which (if any) is more correct? Why?

2. You go to a convenience store to buy candy and find the owner to be rather odd. He allows you to buy pieces in multiples of four, and to buy four, you need \$0.23. He only allows you to do this by using 3 pennies and 2 dimes. You have a bunch of pennies and dimes, and instead of counting them, you decide to weigh them. You have 636.3 g of pennies, and each penny weighs 3.03 g. Each dime weighs 2.29 g. Each piece of candy weighs 10.23 g.

a. How many pennies do you have?
b. How many dimes do you need to buy as much candy as possible?
c. How much should all these dimes weigh?
d. How many pieces of candy could you buy? (number of dimes from part b)
e. How much would this candy weigh?
f. How many pieces of candy could you buy with twice as many dimes?

3. When a marble is dropped into a beaker of water, it sinks to the bottom. Which of the following is the best explanation?

a. The surface area of the marble is not large enough to be held up by the surface tension of the water.
b. The mass of the marble is greater than that of the water.
c. The marble weighs more than an equivalent volume of the water.
d. The force from dropping the marble breaks the surface tension of the water.
e. The marble has greater mass and volume than the water.

Justify your choice, and for choices you did not pick, explain what is wrong about them.

4. You have two beakers, one filled to the 100-mL mark with sugar (the sugar has a mass of 180.0 g) and the other filled to the 100-mL mark with water (the water has a mass of 100.0 g). You pour all the sugar and all the water together in a bigger beaker and stir until the sugar is completely dissolved.

a. Which of the following is true about the mass of the solution? Explain.

i. It is much greater than 280.0 g.
ii. It is somewhat greater than 280.0 g.
iii. It is exactly 280.0 g.
iv. It is somewhat less than 280.0 g.
v. It is much less than 280.0 g.

b. Which of the following is true about the volume of the solution? Explain.

i. It is much greater than 200.0 mL.
ii. It is somewhat greater than 200.0 mL.
iii. It is exactly 200.0 mL.
iv. It is somewhat less than 200.0 mL.
v. It is much less than 200.0 mL.

5. You may have noticed that when water boils, you can see bubbles that rise to the surface of the water.

a. What is inside these bubbles?

i. air
ii. hydrogen and oxygen gas
iii. oxygen gas
iv. water vapor
v. carbon dioxide gas

b. Is the boiling of water a chemical or physical change? Explain.

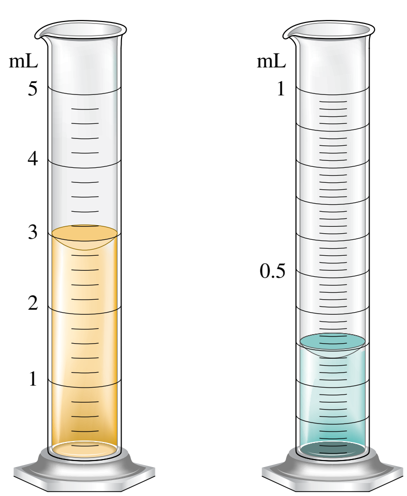
6. If you place a glass rod over a burning candle, the glass appears to turn black. What is happening to each of the following (physical change, chemical change, both, or neither) as the candle burns? Explain each answer.

a. the wax **b.** the wick **c.** the glass rod

7. Which characteristics of a solid, a liquid, and a gas are exhibited by each of the following substances? How would you classify each substance?

a. a bowl of pudding **b.** a bucketful of sand

8. You have water in each graduated cylinder shown:



You then add both samples to a beaker. How would you write the number describing the total volume? What limits the precision of this number?

9. Paracelsus, a sixteenth-century alchemist and healer, adopted as his slogan: "The patients are your textbook, the sickbed is your study." Is this view consistent with using the scientific method?

10. What is wrong with the following statement?

"The results of the experiment do not agree with the theory. Something must be wrong with the experiment."

11. Why is it incorrect to say that the results of a measurement were accurate but not precise?

12. What data would you need to estimate the money you would spend on gasoline to drive your car from New York to Chicago? Provide estimates of values and a sample calculation.

13. Sketch two pieces of glassware: one that can measure volume to the thousandths place and one that can measure volume only to the ones place.

14. You have a 1.0-cm^3 sample of lead and a 1.0-cm^3 sample of glass. You drop each in separate beakers of water. How do the volumes of water displaced by each sample compare? Explain.

15. Sketch a magnified view (showing atoms/molecules) of each of the following and explain:

a. a heterogeneous mixture of two different compounds
b. a homogeneous mixture of an element and a compound

16. You are driving 65 mi/h and take your eyes off the road for "just a second." What distance (in feet) do you travel in this time?

A blue question or exercise number indicates that the answer to that question or exercise appears at the back of this book and a solution appears in the Solutions Guide.

Questions

17. The difference between a *law* and a *theory* is the difference between *what* and *why*. Explain.

18. Explain the fundamental steps of the scientific method.

19. A measurement is a quantitative observation involving both a number and a unit. What is a qualitative observation? What are the SI units for mass, length, and volume? What is the assumed uncertainty in a number (unless stated otherwise)? The uncertainty of a measurement depends on the precision of the measuring device. Explain.

20. To determine the volume of a cube, a student measured one of the dimensions of the cube several times. If the true dimension of the cube is 10.62 cm, give an example of four sets of measurements that would illustrate the following.

a. imprecise and inaccurate data
b. precise but inaccurate data
c. precise and accurate data

Give a possible explanation as to why data can be imprecise or inaccurate. What is wrong with saying a set of measurements is imprecise but accurate?

21. What are significant figures? Show how to indicate the number one thousand to 1 significant figure, 2 significant figures, 3 significant figures, and 4 significant figures. Why is the answer, to the correct number of significant figures, not 1.0 for the following calculation?

$$\frac{1.5 - 1.0}{0.50} =$$

22. What is the volume per unit mass equal to? What unit conversion would the volume per unit mass be useful for?

23. When the temperature in degrees Fahrenheit (T_F) is plotted versus the temperature in degrees Celsius (T_C), a straight line plot results. A straight line plot also results when T_C is plotted versus T_K (the temperature in degrees Kelvin). Reference Appendix A1.3 and determine the slope and y-intercept of each of these two plots.

24. Give four examples illustrating each of the following terms.

a. homogeneous mixture
b. heterogeneous mixture
c. compound
d. element
e. physical change
f. chemical change

Exercises

In this section similar exercises are paired.

Significant Figures and Unit Conversions

25. Which of the following are exact numbers?

a. There are *100* cm in 1 m.
b. One meter equals *1.094* yard.
c. We can use the equation

$$°F = \tfrac{9}{5}°C + 32$$

to convert from Celsius to Fahrenheit temperature. Are the numbers $\frac{9}{5}$ and *32* exact or inexact?
d. $\pi = 3.1415927$.

26. Indicate the number of significant figures in each of the following:

a. This book contains more than 1000 pages.
b. A mile is about 5300 ft.
c. A liter is equivalent to 1.059 qt.
d. The population of the United States is approaching 3.0×10^2 million.
e. A kilogram is 1000 g.
f. The Boeing 747 cruises at around 600 mi/h.

27. How many significant figures are there in each of the following values?

a. 6.07×10^{-15}
b. 0.003840
c. 17.00
d. 8×10^8
e. 463.8052
f. 300
g. 301
h. 300.

28. How many significant figures are in each of the following?

a. 100
b. 1.0×10^2
c. 1.00×10^3
d. 100.
e. 0.0048
f. 0.00480
g. 4.80×10^{-3}
h. 4.800×10^{-3}

29. Round off each of the following numbers to the indicated number of significant digits and write the answer in standard scientific notation.

a. 0.00034159 to three digits
b. 103.351×10^2 to four digits
c. 17.9915 to five digits
d. 3.365×10^5 to three digits

30. Use exponential notation to express the number 480 to

a. one significant figure
b. two significant figures
c. three significant figures
d. four significant figures

31. Evaluate each of the following and write the answer to the appropriate number of significant figures.

a. $212.2 + 26.7 + 402.09$
b. $1.0028 + 0.221 + 0.10337$
c. $52.331 + 26.01 - 0.9981$
d. $2.01 \times 10^2 + 3.014 \times 10^3$
e. $7.255 - 6.8350$

32. Perform the following mathematical operations, and express each result to the correct number of significant figures.

a. $\dfrac{0.102 \times 0.0821 \times 273}{1.01}$
b. $0.14 \times 6.022 \times 10^{23}$
c. $4.0 \times 10^4 \times 5.021 \times 10^{-3} \times 7.34993 \times 10^2$
d. $\dfrac{2.00 \times 10^6}{3.00 \times 10^{-7}}$

33. Perform the following mathematical operations and express the result to the correct number of significant figures.

a. $\dfrac{2.526}{3.1} + \dfrac{0.470}{0.623} + \dfrac{80.705}{0.4326}$
b. $(6.404 \times 2.91)/(18.7 - 17.1)$
c. $6.071 \times 10^{-5} - 8.2 \times 10^{-6} - 0.521 \times 10^{-4}$
d. $(3.8 \times 10^{-12} + 4.0 \times 10^{-13})/(4 \times 10^{12} + 6.3 \times 10^{13})$
e. $\dfrac{9.5 + 4.1 + 2.8 + 3.175}{4}$

(Assume that this operation is taking the average of four numbers. Thus 4 in the denominator is exact.)

f. $\dfrac{8.925 - 8.905}{8.925} \times 100$

(This type of calculation is done many times in calculating a percentage error. Assume that this example is such a calculation; thus 100 can be considered to be an exact number.)

34. Perform the following mathematical operations, and express the result to the correct number of significant figures.

a. $6.022 \times 10^{23} \times 1.05 \times 10^2$
b. $\dfrac{6.6262 \times 10^{-34} \times 2.998 \times 10^8}{2.54 \times 10^{-9}}$
c. $1.285 \times 10^{-2} + 1.24 \times 10^{-3} + 1.879 \times 10^{-1}$
d. $1.285 \times 10^{-2} - 1.24 \times 10^{-3}$
e. $\dfrac{(1.00866 - 1.00728)}{6.02205 \times 10^{23}}$
f. $\dfrac{9.875 \times 10^2 - 9.795 \times 10^2}{9.875 \times 10^2} \times 100$ (100 is exact)
g. $\dfrac{9.42 \times 10^2 + 8.234 \times 10^2 + 1.625 \times 10^3}{3}$ (3 is exact)

35. Perform each of the following conversions.

a. 8.43 cm to millimeters
b. 2.41×10^2 cm to meters
c. 294.5 nm to centimeters
d. 1.445×10^4 m to kilometers

e. 235.3 m to millimeters
f. 903.3 nm to micrometers

36. a. How many kilograms are in one teragram?
b. How many nanometers are in 6.50×10^2 terameters?
c. How many kilograms are in 25 femtograms?
d. How many liters are in 8.0 cubic decimeters?
e. How many microliters are in one milliliter?
f. How many picograms are in one microgram?

37. Perform the following unit conversions.
a. Congratulations! You and your spouse are the proud parents of a new baby, born while you are studying in a country that uses the metric system. The nurse has informed you that the baby weighs 3.91 kg and measures 51.4 cm. Convert your baby's weight to pounds and ounces and her length to inches (rounded to the nearest quarter inch).
b. The circumference of the earth is 25,000 mi at the equator. What is the circumference in kilometers? in meters?
c. A rectangular solid measures 1.0 m by 5.6 cm by 2.1 dm. Express its volume in cubic meters, liters, cubic inches, and cubic feet.

38. Perform the following unit conversions.
a. 908 oz to kilograms
b. 12.8 L to gallons
c. 125 mL to quarts
d. 2.89 gal to milliliters
e. 4.48 lb to grams
f. 550 mL to quarts

39. Use the following exact conversion factors to perform the stated calculations:

$$5\tfrac{1}{2} \text{ yards} = 1 \text{ rod}$$
$$40 \text{ rods} = 1 \text{ furlong}$$
$$8 \text{ furlongs} = 1 \text{ mile}$$

a. The Kentucky Derby race is 1.25 miles. How long is the race in rods, furlongs, meters, and kilometers?
b. A marathon race is 26 miles, 385 yards. What is this distance in rods, furlongs, meters, and kilometers?

40. Although the preferred SI unit of area is the square meter, land is often measured in the metric system in hectares (ha). One hectare is equal to 10,000 m^2. In the English system, land is often measured in acres (1 acre = 160 rod^2). Use the exact conversions and those given in Exercise 39 to calculate the following.
a. 1 ha = ________ km^2.
b. The area of a 5.5-acre plot of land in hectares, square meters, and square kilometers.
c. A lot with dimensions 120 ft by 75 ft is to be sold for $6500. What is the price per acre? What is the price per hectare?

41. Precious metals and gems are measured in troy weights in the English system:

$$24 \text{ grains} = 1 \text{ pennyweight (exact)}$$
$$20 \text{ pennyweight} = 1 \text{ troy ounce (exact)}$$
$$12 \text{ troy ounces} = 1 \text{ troy pound (exact)}$$
$$1 \text{ grain} = 0.0648 \text{ gram}$$
$$1 \text{ carat} = 0.200 \text{ gram}$$

a. The most common English unit of mass is the pound avoirdupois. What is one troy pound in kilograms and in pounds?
b. What is the mass of a troy ounce of gold in grams and in carats?
c. The density of gold is 19.3 g/cm^3. What is the volume of a troy pound of gold?

42. Apothecaries (druggists) use the following set of measures in the English system:

$$20 \text{ grains ap} = 1 \text{ scruple (exact)}$$
$$3 \text{ scruples} = 1 \text{ dram ap (exact)}$$
$$8 \text{ dram ap} = 1 \text{ oz ap (exact)}$$
$$1 \text{ dram ap} = 3.888 \text{ g}$$

a. Is an apothecary grain the same as a troy grain? (See Exercise 41.)
b. 1 oz ap = ________ oz troy.
c. An aspirin tablet contains 5.00×10^2 mg of active ingredient. What mass in grains ap of active ingredient does it contain? What mass in scruples?
d. What is the mass of 1 scruple in grams?

43. Science fiction often uses nautical analogies to describe space travel. If the starship *U.S.S. Enterprise* is traveling at warp factor 1.71, what is its speed in knots and in miles per hour? (Warp 1.71 = 5.00 times the speed of light; speed of light = 3.00×10^8 m/s; 1 knot = 2000 yd/h, exactly.)

44. The world record for the hundred meter dash is 9.77 s. What is the corresponding average speed in units of m/s, km/h, ft/s, and mi/h? At this speed, how long would it take to run 1.00×10^2 yards?

45. Would a car traveling at a constant speed of 65 km/h violate a 40. mi/h speed limit?

46. You pass a road sign saying "New York 112 km." If you drive at a constant speed of 65 mi/h, how long should it take you to reach New York? If your car gets 28 miles to the gallon, how many liters of gasoline are necessary to travel 112 km?

47. If you put 8.21 gallons of gas in your car and it cost you a total of $17.25, what is the cost of gas per liter in Canadian dollars? Assume 0.82 dollar U.S. = 1.00 dollar Canadian.

48. A children's pain relief elixir contains 80. mg acetaminophen per 0.50 teaspoon. The dosage recommended for a child who weighs between 24 and 35 lb is 1.5 teaspoons. What is the range of acetaminophen dosages, expressed in mg acetaminophen/kg body weight, for children who weigh between 24 and 35 lb?

Temperature

49. Convert the following Fahrenheit temperatures to the Celsius and Kelvin scales.
a. −459°F, an extremely low temperature
b. −40.°F, the answer to a trivia question
c. 68°F, room temperature
d. 7×10^7 °F, temperature required to initiate fusion reactions in the sun

50. A thermometer gives a reading of 96.1°F ± 0.2°F. What is the temperature in °C? What is the uncertainty?

51. Convert the following Celsius temperatures to Kelvin and to Fahrenheit degrees.
a. the temperature of someone with a fever, 39.2°C
b. a cold wintery day, −25°C
c. the lowest possible temperature, −273°C
d. the melting-point temperature of sodium chloride, 801°C

52. Convert the following Kelvin temperatures to Celsius and Fahrenheit degrees.
   a. the temperature that registers the same value on both the Fahrenheit and Celsius scales, 233 K
   b. the boiling point of helium, 4 K
   c. the temperature at which many chemical quantities are determined, 298 K
   d. the melting point of tungsten, 3680 K

Density

53. A material will float on the surface of a liquid if the material has a density less than that of the liquid. Given that the density of water is approximately 1.0 g/mL, will a block of material having a volume of 1.2×10^4 in^3 and weighing 350 lb float or sink when placed in a reservoir of water?

54. For a material to float on the surface of water, the material must have a density less than that of water (1.0 g/mL) and must not react with the water or dissolve in it. A spherical ball has a radius of 0.50 cm and weighs 2.0 g. Will this ball float or sink when placed in water? (*Note:* Volume of a sphere $= \frac{4}{3}\pi r^3$.)

55. A star is estimated to have a mass of 2×10^{36} kg. Assuming it to be a sphere of average radius 7.0×10^5 km, calculate the average density of the star in units of grams per cubic centimeter.

56. A rectangular block has dimensions 2.9 cm $\times$ 3.5 cm $\times$ 10.0 cm. The mass of the block is 615.0 g. What are the volume and density of the block?

57. Diamonds are measured in carats, and 1 carat = 0.200 g. The density of diamond is 3.51 g/cm^3. What is the volume of a 5.0-carat diamond?

58. The volume of a diamond is found to be 2.8 mL. What is the mass of the diamond in carats? (See Exercise 57.)

59. A sample containing 33.42 g of metal pellets is poured into a graduated cylinder initially containing 12.7 mL of water, causing the water level in the cylinder to rise to 21.6 mL. Calculate the density of the metal.

60. The density of pure silver is 10.5 g/cm^3 at 20°C. If 5.25 g of pure silver pellets is added to a graduated cylinder containing 11.2 mL of water, to what volume level will the water in the cylinder rise?

61. In each of the following pairs, which has the greater mass? (See Table 1.5.)
   a. 1.0 kg of feathers or 1.0 kg of lead
   b. 1.0 mL of mercury or 1.0 mL of water
   c. 19.3 mL of water or 1.00 mL of gold
   d. 75 mL of copper or 1.0 L of benzene

62. Mercury poisoning is a debilitating disease that is often fatal. In the human body, mercury reacts with essential enzymes leading to irreversible inactivity of these enzymes. If the amount of mercury in a polluted lake is 0.4 μg Hg/mL, what is the total mass in kilograms of mercury in the lake? (The lake has a surface area of 100 mi^2 and an average depth of 20 ft.)

63. In each of the following pairs, which has the greater volume?
   a. 1.0 kg of feathers or 1.0 kg of lead
   b. 100 g of gold or 100 g of water
   c. 1.0 L of copper or 1.0 L of mercury

64. Using Table 1.5, calculate the volume of 25.0 g of each of the following substances at 1 atm.
   a. hydrogen gas
   b. water
   c. iron

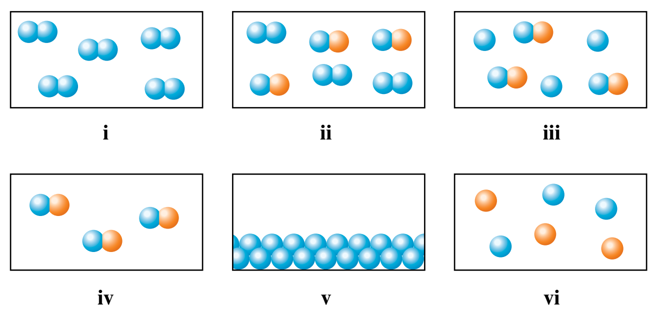
   Chapter 5 discusses the properties of gases. One property unique to gases is that they contain mostly empty space. Explain using the results of your calculations.

65. The density of osmium (the densest metal) is 22.57 g/cm^3. If a 1.00-kg rectangular block of osmium has two dimensions of 4.00 cm $\times$ 4.00 cm, calculate the third dimension of the block.

66. A copper wire (density = 8.96 g/cm^3) has a diameter of 0.25 mm. If a sample of this copper wire has a mass of 22 g, how long is the wire?

Classification and Separation of Matter

67. Match each description below with the following microscopic pictures. More than one picture may fit each description. A picture may be used more than once or not used at all.



   a. a gaseous compound
   b. a mixture of two gaseous elements
   c. a solid element
   d. a mixture of a gaseous element and a gaseous compound

68. Define the following terms: solid, liquid, gas, pure substance, element, compound, homogeneous mixture, heterogeneous mixture, solution, chemical change, physical change.

69. What is the difference between homogeneous and heterogeneous matter? Classify each of the following as homogeneous or heterogeneous.
   a. a door
   b. the air you breathe
   c. a cup of coffee (black)
   d. the water you drink
   e. salsa
   f. your lab partner

70. Classify each of the following as a mixture or a pure substance.
   a. water
   b. blood
   c. the oceans
   d. iron
   e. brass
   f. uranium
   g. wine
   h. leather
   i. table salt (NaCl)

   Of the pure substances, which are elements and which are compounds?

71. Classify following as physical or chemical changes.

a. Moth balls gradually vaporize in a closet.

b. Hydrofluoric acid attacks glass, and is used to etch calibration marks on glass laboratory utensils.

c. A French chef making a sauce with brandy is able to burn off the alcohol from the brandy, leaving just the brandy flavoring.

d. Chemistry majors sometimes get holes in the cotton jeans they wear to lab because of acid spills.

72. The properties of a mixture are typically averages of the properties of its components. The properties of a compound may differ dramatically from the properties of the elements that combine to produce the compound. For each process described below, state whether the material being discussed is most likely a mixture or a compound, and state whether the process is a chemical change or a physical change.

a. An orange liquid is distilled, resulting in the collection of a yellow liquid and a red solid.

b. A colorless, crystalline solid is decomposed, yielding a pale yellow-green gas and a soft, shiny metal.

c. A cup of tea becomes sweeter as sugar is added to it.

Additional Exercises

73. For a pharmacist dispensing pills or capsules, it is often easier to weigh the medication to be dispensed rather than to count the individual pills. If a single antibiotic capsule weighs 0.65 g, and a pharmacist weighs out 15.6 g of capsules, how many capsules have been dispensed?

74. In Shakespeare's *Richard III,* the First Murderer says:

"Take that, and that! [*Stabs Clarence*]
If that is not enough, I'll drown you in a malmsey butt within!"

Given that 1 butt = 126 gal, in how many liters of malmsey (a foul brew similar to mead) was the unfortunate Clarence about to be drowned?

75. The contents of one 40. lb bag of topsoil will cover 10. square feet of ground to a depth of 1.0 inch. What number of bags are needed to cover a plot that measures 200. by 300. m to a depth of 4.0 cm?

76. In the opening scenes of the movie *Raiders of the Lost Ark,* Indiana Jones tries to remove a gold idol from a booby-trapped pedestal. He replaces the idol with a bag of sand of approximately equal volume. (Density of gold = 19.32 g/cm^3; density of sand $\approx$ 2 g/cm^3.)

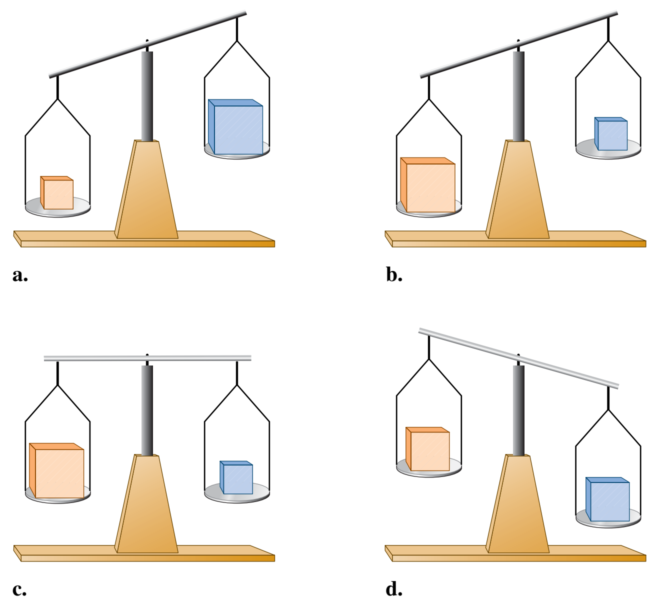
a. Did he have a reasonable chance of not activating the mass-sensitive booby trap?

b. In a later scene he and an unscrupulous guide play catch with the idol. Assume that the volume of the idol is about 1.0 L. If it were solid gold, what mass would the idol have? Is playing catch with it plausible?

77. A column of liquid is found to expand linearly on heating 5.25 cm for a 10.0°F rise in temperature. If the initial temperature of the liquid is 98.6°F, what will the final temperature be in °C if the liquid has expanded by 18.5 cm?

78. A 25.00-g sample of a solid is placed in a graduated cylinder and then the cylinder is filled to the 50.0 mL mark with benzene. The mass of benzene and solid together is 58.80 g. Assuming that the solid is insoluble in benzene and that the density of benzene is 0.880 g/cm^3, calculate the density of the solid.

79. For each of the following, decide which block is more dense: the orange block, the blue block, or it cannot be determined. Explain your answers.



80. According to the *Official Rules of Baseball,* a baseball must have a circumference not more than 9.25 in or less than 9.00 in and a mass not more than 5.25 oz or less than 5.00 oz. What range of densities can a baseball be expected to have? Express this range as a single number with an accompanying uncertainty limit.

81. The density of an irregularly shaped object was determined as follows. The mass of the object was found to be 28.90 g $\pm$ 0.03 g. A graduated cylinder was partially filled with water. The reading of the level of the water was 6.4 cm^3 $\pm$ 0.1 cm^3. The object was dropped in the cylinder, and the level of the water rose to 9.8 cm^3 $\pm$ 0.1 cm^3. What is the density of the object with appropriate error limits? (See Appendix 1.5.)

Challenge Problems

82. Draw a picture showing the markings (graduations) on glassware that would allow you to make each of the following volume measurements of water and explain your answers (the numbers given are as precise as possible).

a. 128.7 mL **b.** 18 mL **c.** 23.45 mL

If you made the measurements of three samples of water and then poured all of the water together in one container, what total volume of water should you report? Support your answer.

83. Many times errors are expressed in terms of percentage. The percent error is the absolute value of the difference of the true value and the experimental value, divided by the true value, and multiplied by 100.

$$\text{Percent error} = \frac{|\text{true value} - \text{experimental value}|}{\text{true value}} \times 100$$

Calculate the percent error for the following measurements.

a. The density of an aluminum block determined in an experiment was 2.64 g/cm^3. (True value 2.70 g/cm^3.)

b. The experimental determination of iron in iron ore was 16.48%. (True value 16.12%.)

c. A balance measured the mass of a 1.000-g standard as 0.9981 g.

84. A person weighed 15 pennies on a balance and recorded the following masses:

3.112 g	3.109 g	3.059 g
2.467 g	3.079 g	2.518 g
3.129 g	2.545 g	3.050 g
3.053 g	3.054 g	3.072 g
3.081 g	3.131 g	3.064 g

Curious about the results, he looked at the dates on each penny. Two of the light pennies were minted in 1983 and one in 1982. The dates on the 12 heavier pennies ranged from 1970 to 1982. Two of the 12 heavier pennies were minted in 1982.

a. Do you think the Bureau of the Mint changed the way it made pennies? Explain.

b. The person calculated the average mass of the 12 heavy pennies. He expressed this average as 3.0828 g ± 0.0482 g. What is wrong with the numbers in this result, and how should the value be expressed?

85. On October 21, 1982, the Bureau of the Mint changed the composition of pennies (see Exercise 84). Instead of an alloy of 95% Cu and 5% Zn by mass, a core of 99.2% Zn and 0.8% Cu with a thin shell of copper was adopted. The overall composition of the new penny was 97.6% Zn and 2.4% Cu by mass. Does this account for the difference in mass among the pennies in Exercise 84? Assume the volume of the individual metals that make up each penny can be added together to give the overall volume of the penny, and assume each penny is the same size. (Density of Cu = 8.96 g/cm^3; density of Zn = 7.14 g/cm^3.)

86. Ethylene glycol is the main component in automobile antifreeze. To monitor the temperature of an auto cooling system, you intend to use a meter that reads from 0 to 100. You devise a new temperature scale based on the approximate melting and boiling points of a typical antifreeze solution (−45°C and 115°C). You wish these points to correspond to 0°A and 100°A, respectively.

a. Derive an expression for converting between °A and °C.

b. Derive an expression for converting between °F and °A.

c. At what temperature would your thermometer and a Celsius thermometer give the same numerical reading?

d. Your thermometer reads 86°A. What is the temperature in °C and in °F?

e. What is a temperature of 45°C in °A?

87. Sterling silver is a solid solution of silver and copper. If a piece of a sterling silver necklace has a mass of 105.0 g and a volume of 10.12 mL, calculate the mass percent of copper in the piece of necklace. Assume that the volume of silver present plus the volume of copper present equals the total volume. Refer to Table 1.5.

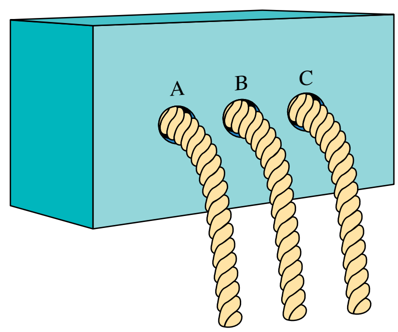
$$\text{Mass percent of copper} = \frac{\text{mass of copper}}{\text{total mass}} \times 100$$

88. Use molecular-level (microscopic) drawings for each of the following.

a. Show the differences between a gaseous mixture that is a homogeneous mixture of two different compounds, and a gaseous mixture that is a homogeneous mixture of a compound and an element.

b. Show the differences among a gaseous element, a liquid element, and a solid element.

89. Confronted with the box shown in the diagram, you wish to discover something about its internal workings. You have no tools and cannot open the box. You pull on rope B, and it moves rather freely. When you pull on rope A, rope C appears to be pulled slightly into the box. When you pull on rope C, rope A almost disappears into the box.*



a. Based on these observations, construct a model for the interior mechanism of the box.

b. What further experiments could you do to refine your model?

90. An experiment was performed in which an empty 100-mL graduated cylinder was weighed. It was weighed once again after it had been filled to the 10.0-mL mark with dry sand. A 10-mL pipet was used to transfer 10.00 mL of methanol to the cylinder. The sand–methanol mixture was stirred until bubbles no longer emerged from the mixture and the sand looked uniformly wet. The cylinder was then weighed again. Use the data obtained from this experiment (and displayed at the end of this problem) to find the density of the dry sand, the density of methanol, and the density of sand particles. Does the bubbling that occurs when the methanol is added to the dry sand indicate that the sand and methanol are reacting?

Mass of cylinder plus wet sand	45.2613 g
Mass of cylinder plus dry sand	37.3488 g
Mass of empty cylinder	22.8317 g
Volume of dry sand	10.0 mL
Volume of sand + methanol	17.6 mL
Volume of methanol	10.00 mL

*From Yoder, Suydam, and Snavely, *Chemistry* (New York: Harcourt Brace Jovanovich, 1975), pp. 9–11.

Integrative Problems

These problems require the integration of multiple concepts to find the solutions.

91. The U.S. trade deficit at the beginning of 2005 was \$475,000,000. If the wealthiest 1.00 percent of the U.S. population (297,000,000) contributed an equal amount of money to bring the trade deficit to \$0, how many dollars would each person contribute? If one of these people were to pay their share in nickels only, how many nickels are needed? Another person living abroad at the time decides to pay in pounds sterling (£). How many pounds sterling does this person contribute (assume a conversion rate of 1 £ = \$ 1.869)?

92. The density of osmium is reported by one source to be 22610 kg/m^3. What is this density in g/cm^3? What is the mass of a block of osmium measuring 10.0 cm $\times$ 8.0 cm $\times$ 9.0 cm?

93. At the Amundsen-Scott South Pole base station in Antarctica, when the temperature is -100.0°F, researchers who live there can join the "300 Club" by stepping into a sauna heated to 200.0°F then quickly running outside and around the pole that marks the South Pole. What are these temperatures in °C? What are these temperatures in K? If you measured the temperatures only in °C and K, can you become a member of the "300 Club" (that is, is there a 300.-degree difference between the temperature extremes when measured in °C and K?)

Marathon Problem*

This problem is designed to incorporate several concepts and techniques into one situation. Marathon Problems can be used in class by groups of students to help facilitate problem-solving skills.

94. A cylindrical bar of gold that is 1.5 in high and 0.25 in in diameter has a mass of 23.1984 g, as determined on an analytical balance. An empty graduated cylinder is weighed on a triple-beam balance and has a mass of 73.47 g. After pouring a small amount of a liquid into the graduated cylinder, the mass is 79.16 g. When the gold cylinder is placed in the graduated cylinder (the liquid covers the top of the gold cylinder), the volume indicated on the graduated cylinder is 8.5 mL. Assume that the temperature of the gold bar and the liquid are 86°F. If the density of the liquid decreases by 1.0% for each 10.°C rise in temperature (over the range 0 to 50°C), determine

a. the density of the gold at 86°F.
b. the density of the liquid at 40.°F.

Note: Parts a and b can be answered independently.

Get help understanding core concepts and visualizing molecular-level interactions, and practice problem solving, by visiting the Online Study Center at **college.hmco.com/PIC/zumdahl7e.**

2 Atoms, Molecules, and Ions

Contents

A worker in Thailand piles up salt crystals.

Where does one start in learning chemistry? Clearly we must consider some essential vocabulary and something about the origins of the science before we can proceed very far. Thus, while Chapter 1 provided background on the fundamental ideas and procedures of science in general, Chapter 2 covers the specific chemical background necessary for understanding the material in the next few chapters. The coverage of these topics is necessarily brief at this point. We will develop these ideas more fully as it becomes appropriate to do so. A major goal of this chapter is to present the systems for naming chemical compounds to provide you with the vocabulary necessary to understand this book and to pursue your laboratory studies.

Because chemistry is concerned first and foremost with chemical changes, we will proceed as quickly as possible to a study of chemical reactions (Chapters 3 and 4). However, before we can discuss reactions, we must consider some fundamental ideas about atoms and how they combine.

2.1 The Early History of Chemistry

Chemistry has been important since ancient times. The processing of natural ores to produce metals for ornaments and weapons and the use of embalming fluids are just two applications of chemical phenomena that were utilized prior to 1000 B.C.

The Greeks were the first to try to explain why chemical changes occur. By about 400 B.C. they had proposed that all matter was composed of four fundamental substances: fire, earth, water, and air. The Greeks also considered the question of whether matter is continuous, and thus infinitely divisible into smaller pieces, or composed of small, indivisible particles. Supporters of the latter position were Demokritos* of Abdera (*c.* 460–*c.* 370 B.C.) and Leucippos, who used the term *atomos* (which later became *atoms*) to describe these ultimate particles. However, because the Greeks had no experiments to test their ideas, no definitive conclusion could be reached about the divisibility of matter.

The next 2000 years of chemical history were dominated by a pseudoscience called *alchemy.* Some alchemists were mystics and fakes who were obsessed with the idea of turning cheap metals into gold. However, many alchemists were serious scientists, and this period saw important advances: The alchemists discovered several elements and learned to prepare the mineral acids.

The foundations of modern chemistry were laid in the sixteenth century with the development of systematic metallurgy (extraction of metals from ores) by a German, Georg Bauer (1494–1555), and the medicinal application of minerals by a Swiss alchemist/physician known as Paracelsus (full name: Philippus Theophrastus Bombastus von Hohenheim [1493–1541]).

The first "chemist" to perform truly quantitative experiments was Robert Boyle (1627–1691), who carefully measured the relationship between the pressure and volume of air. When Boyle published his book *The Skeptical Chymist* in 1661, the quantitative sciences of physics and chemistry were born. In addition to his results on the quantitative behavior of gases, Boyle's other major contribution to chemistry consisted of his ideas about the chemical elements. Boyle held no preconceived notion about the number of elements. In his view, a substance was an element unless it could be broken down into two or more simpler substances. As Boyle's experimental definition of an element became generally accepted, the list of known elements began to grow, and the Greek system of

*Democritus is an alternate spelling.

CHEMICAL IMPACT

There's Gold in Them There Plants!

Gold has always held a strong allure. For example, the alchemists were obsessed with finding a way to transform cheap metals into gold. Also, when gold was discovered in California in 1849, a frantic rush occurred to that area and other areas in the west. Although gold is still valuable, most of the high-grade gold ores have been exhausted. This leaves the low-grade ores—ores with low concentrations of gold that are expensive to process relative to the amount of gold finally obtained.

Now two scientists have come across a novel way to concentrate the gold from low-grade ores. Christopher Anderson and Robert Brooks of Massey University in Palmerston North, New Zealand, have found plants that accumulate gold atoms as they grow in soil containing gold ore [*Nature* 395 (1998): 553]. The plants brassica (of the mustard family) and chicory seem especially effective as botanical "gold miners." When these plants are dried and burned (after having grown in gold-rich soil), the resulting ash contains approximately 150 ppm (parts per million) of gold. (1 ppm gold represents 1 g of gold in 10^6 g of sample.)

The New Zealand scientists were able to double the amount of gold taken from the soil by the plants by treating the soil with ammonium thiocyanate (NH_4SCN). The thiocyanate, which reacts with the gold, making it more available to the plants, subsequently breaks down in the soil and therefore poses no environmental hazard.

Thus plants seem to hold great promise as gold miners. They are efficient and reliable and will never go on strike.

This plant from the mustard family is a newly discovered source of gold.

four elements finally died. Although Boyle was an excellent scientist, he was not always right. For example, he clung to the alchemists' views that metals were not true elements and that a way would eventually be found to change one metal into another.

The phenomenon of combustion evoked intense interest in the seventeenth and eighteenth centuries. The German chemist Georg Stahl (1660–1734) suggested that a substance he called "phlogiston" flowed out of the burning material. Stahl postulated that a substance burning in a closed container eventually stopped burning because the air in the container became saturated with phlogiston. Oxygen gas, discovered by Joseph Priestley (1733–1804),* an English clergyman and scientist (Fig. 2.1), was found to support vigorous combustion and was thus supposed to be low in phlogiston. In fact, oxygen was originally called "dephlogisticated air."

FIGURE 2.1
The Priestley Medal is the highest honor given by the American Chemical Society. It is named for Joseph Priestley, who was born in England on March 13, 1733. He performed many important scientific experiments, among them the discovery that a gas later identified as carbon dioxide could be dissolved in water to produce *seltzer*. Also, as a result of meeting Benjamin Franklin in London in 1766, Priestley became interested in electricity and was the first to observe that graphite was an electrical conductor. However, his greatest discovery occurred in 1774 when he isolated oxygen by heating mercuric oxide.

Because of his nonconformist political views, Priestley was forced to leave England. He died in the United States in 1804.

*Oxygen gas was actually first observed by the Swedish chemist Karl W. Scheele (1742–1786), but because his results were published after Priestley's, the latter is commonly credited with the discovery of oxygen.

2.2 Fundamental Chemical Laws

By the late eighteenth century, combustion had been studied extensively; the gases carbon dioxide, nitrogen, hydrogen, and oxygen had been discovered; and the list of elements continued to grow. However, it was Antoine Lavoisier (1743–1794), a French chemist (Fig. 2.2), who finally explained the true nature of combustion, thus clearing the way for the tremendous progress that was made near the end of the eighteenth century. Lavoisier, like Boyle, regarded measurement as the essential operation of chemistry. His experiments, in which he carefully weighed the reactants and products of various reactions, suggested that *mass is neither created nor destroyed.* Lavoisier's verification of this **law of conservation of mass** was the basis for the developments in chemistry in the nineteenth century.

Oxygen is from the French *oxygène,* meaning "generator of acid," because it was initially considered to be an integral part of all acids.

Lavoisier's quantitative experiments showed that combustion involved oxygen (which Lavoisier named), not phlogiston. He also discovered that life was supported by a process that also involved oxygen and was similar in many ways to combustion. In 1789 Lavoisier published the first modern chemistry textbook, *Elementary Treatise on Chemistry,* in which he presented a unified picture of the chemical knowledge assembled up to that time. Unfortunately, in the same year the text was published, the French Revolution broke out. Lavoisier, who had been associated with collecting taxes for the government, was executed on the guillotine as an enemy of the people in 1794.

After 1800, chemistry was dominated by scientists who, following Lavoisier's lead, performed careful weighing experiments to study the course of chemical reactions and to determine the composition of various chemical compounds. One of these chemists, a Frenchman, Joseph Proust (1754–1826), showed that *a given compound always contains exactly the same proportion of elements by mass.* For example, Proust found that the substance copper carbonate is always 5.3 parts copper to 4 parts oxygen to 1 part carbon (by mass). The principle of the constant composition of compounds, originally called "Proust's law," is now known as the **law of definite proportion.**

Proust's discovery stimulated John Dalton (1766–1844), an English schoolteacher (Fig. 2.3), to think about atoms as the particles that might compose elements. Dalton reasoned that if elements were composed of tiny individual particles, a given compound should always contain the same combination of these atoms. This concept explained why the same relative masses of elements were always found in a given compound.

But Dalton discovered another principle that convinced him even more of the existence of atoms. He noted, for example, that carbon and oxygen form two different compounds that contain different relative amounts of carbon and oxygen, as shown by the following data:

	Mass of Oxygen That Combines with 1 g of Carbon
Compound I	1.33 g
Compound II	2.66 g

Dalton noted that compound II contains twice as much oxygen per gram of carbon as compound I, a fact that could easily be explained in terms of atoms. Compound I might be CO, and compound II might be CO_2.* This principle, which was found to apply to compounds of other elements as well, became known as the **law of multiple proportions:** *When two elements form a series of compounds, the ratios of the masses of the second element that combine with 1 gram of the first element can always be reduced to small whole numbers.*

To make sure the significance of this observation is clear, in Sample Exercise 2.1 we will consider data for a series of compounds consisting of nitrogen and oxygen.

Sample Exercise 2.1 Illustrating the Law of Multiple Proportions

The following data were collected for several compounds of nitrogen and oxygen:

	Mass of Nitrogen That Combines with 1 g of Oxygen
Compound A	1.750 g
Compound B	0.8750 g
Compound C	0.4375 g

Show how these data illustrate the law of multiple proportions.

Solution

For the law of multiple proportions to hold, the ratios of the masses of nitrogen combining with 1 gram of oxygen in each pair of compounds should be small whole numbers. We therefore compute the ratios as follows:

$$\frac{A}{B} = \frac{1.750}{0.875} = \frac{2}{1}$$
$$\frac{B}{C} = \frac{0.875}{0.4375} = \frac{2}{1}$$
$$\frac{A}{C} = \frac{1.750}{0.4375} = \frac{4}{1}$$

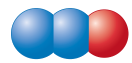
These results support the law of multiple proportions.

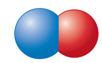
See Exercises 2.27 and 2.28.

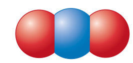
*Subscripts are used to show the numbers of atoms present. The number 1 is understood (not written). The symbols for the elements and the writing of chemical formulas will be illustrated further in Sections 2.6 and 2.7.

The significance of the data in Sample Exercise 2.1 is that compound A contains twice as much nitrogen (N) per gram of oxygen (O) as does compound B and that compound B contains twice as much nitrogen per gram of oxygen as does compound C.

These data can be explained readily if the substances are composed of molecules made up of nitrogen atoms and oxygen atoms. For example, one set of possibilities for compounds A, B, and C is

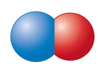
A:  $\frac{N}{O} = \frac{2}{1}$

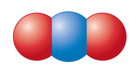
B:  $\frac{N}{O} = \frac{1}{1}$

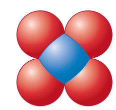
C:  $\frac{N}{O} = \frac{1}{2}$

Now we can see that compound A contains two atoms of N for every atom of O, whereas compound B contains one atom of N per atom of O. That is, compound A contains twice as much nitrogen per given amount of oxygen as does compound B. Similarly, since compound B contains one N per O and compound C contains one N per *two* O's, the nitrogen content of compound C per given amount of oxygen is half that of compound B.

Another set of compounds that fits the data in Sample Exercise 2.1 is

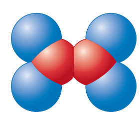
A:  $\frac{N}{O} = \frac{1}{1}$

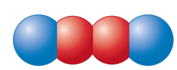
B:  $\frac{N}{O} = \frac{1}{2}$

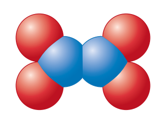
C:  $\frac{N}{O} = \frac{1}{4}$

Verify for yourself that these compounds satisfy the requirements.

Still another set that works is

A:  $\frac{N}{O} = \frac{4}{2}$

B:  $\frac{N}{O} = \frac{2}{2}$

C:  $\frac{N}{O} = \frac{2}{4}$

See if you can come up with still another set of compounds that satisfies the data in Sample Exercise 2.1. How many more possibilities are there?

In fact, an infinite number of other possibilities exists. Dalton could not deduce absolute formulas from the available data on relative masses. However, the data on the composition of compounds in terms of the relative masses of the elements supported his hypothesis that each element consisted of a certain type of atom and that compounds were formed from specific combinations of atoms.

2.3 Dalton's Atomic Theory

In 1808 Dalton published *A New System of Chemical Philosophy,* in which he presented his theory of atoms:

These statements are a modern paraphrase of Dalton's ideas.

1. Each element is made up of tiny particles called atoms.
2. The atoms of a given element are identical; the atoms of different elements are different in some fundamental way or ways.
3. Chemical compounds are formed when atoms of different elements combine with each other. A given compound always has the same relative numbers and types of atoms.
4. Chemical reactions involve reorganization of the atoms—changes in the way they are bound together. The atoms themselves are not changed in a chemical reaction.

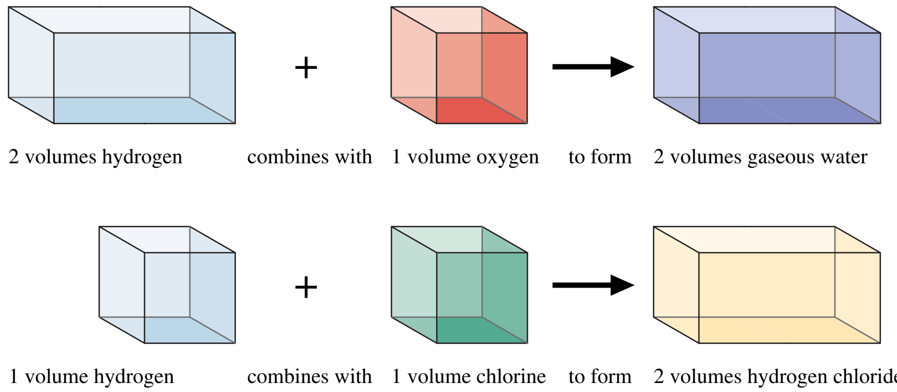


FIGURE 2.4
A representation of some of Gay-Lussac's experimental results on combining gas volumes.

Joseph Louis Gay-Lussac, a French physicist and chemist, was remarkably versatile. Although he is now known primarily for his studies on the combining of volumes of gases, Gay-Lussac was instrumental in the studies of many of the other properties of gases. Some of Gay-Lussac's motivation to learn about gases arose from his passion for ballooning. In fact, he made ascents to heights of over 4 miles to collect air samples, setting altitude records that stood for about 50 years. Gay-Lussac also was the codiscoverer of boron and the developer of a process for manufacturing sulfuric acid. As chief assayer of the French mint, Gay-Lussac developed many techniques for chemical analysis and invented many types of glassware now used routinely in labs. Gay-Lussac spent his last 20 years as a lawmaker in the French government.

It is instructive to consider Dalton's reasoning on the relative masses of the atoms of the various elements. In Dalton's time water was known to be composed of the elements hydrogen and oxygen, with 8 grams of oxygen present for every 1 gram of hydrogen. If the formula for water were OH, an oxygen atom would have to have 8 times the mass of a hydrogen atom. However, if the formula for water were H_2O (two atoms of hydrogen for every oxygen atom), this would mean that each atom of oxygen is 16 times as massive as *each* atom of hydrogen (since the ratio of the mass of one oxygen to that of *two* hydrogens is 8 to 1). Because the formula for water was not then known, Dalton could not specify the relative masses of oxygen and hydrogen unambiguously. To solve the problem, Dalton made a fundamental assumption: He decided that nature would be as simple as possible. This assumption led him to conclude that the formula for water should be OH. He thus assigned hydrogen a mass of 1 and oxygen a mass of 8.

Using similar reasoning for other compounds, Dalton prepared the first table of **atomic masses** (sometimes called **atomic weights** by chemists, since mass is often determined by comparison to a standard mass—a process called *weighing*). Many of the masses were later proved to be wrong because of Dalton's incorrect assumptions about the formulas of certain compounds, but the construction of a table of masses was an important step forward.

Although not recognized as such for many years, the keys to determining absolute formulas for compounds were provided in the experimental work of the French chemist Joseph Gay-Lussac (1778–1850) and by the hypothesis of an Italian chemist named Amadeo Avogadro (1776–1856). In 1809 Gay-Lussac performed experiments in which he measured (under the same conditions of temperature and pressure) the volumes of gases that reacted with each other. For example, Gay-Lussac found that 2 volumes of hydrogen react with 1 volume of oxygen to form 2 volumes of gaseous water and that 1 volume of hydrogen reacts with 1 volume of chlorine to form 2 volumes of hydrogen chloride. These results are represented schematically in Fig. 2.4.

In 1811 Avogadro interpreted these results by proposing that *at the same temperature and pressure, equal volumes of different gases contain the same number of particles.* This assumption (called **Avogadro's hypothesis**) makes sense if the distances between the particles in a gas are very great compared with the sizes of the particles. Under these conditions, the volume of a gas is determined by the number of molecules present, not by the size of the individual particles.

If Avogadro's hypothesis is correct, Gay-Lussac's result,

$$\text{2 volumes of hydrogen react with 1 volume of oxygen} \longrightarrow \text{2 volumes of water vapor}$$

can be expressed as follows:

$$\text{2 molecules* of hydrogen react with 1 molecule of oxygen} \longrightarrow \text{2 molecules of water}$$

*A *molecule* is a collection of atoms (see Section 2.6).

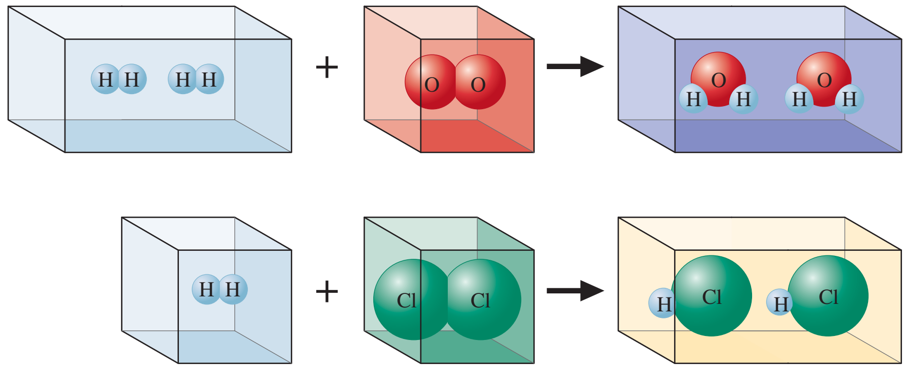


FIGURE 2.5
A representation of combining gases at the molecular level. The spheres represent atoms in the molecules.

These observations can best be explained by assuming that gaseous hydrogen, oxygen, and chlorine are all composed of diatomic (two-atom) molecules: H_2, O_2, and Cl_2, respectively. Gay-Lussac's results can then be represented as shown in Fig. 2.5. (Note that this reasoning suggests that the formula for water is H_2O, not OH as Dalton believed.)

Unfortunately, Avogadro's interpretations were not accepted by most chemists, and a half-century of confusion followed, in which many different assumptions were made about formulas and atomic masses.

During the nineteenth century, painstaking measurements were made of the masses of various elements that combined to form compounds. From these experiments a list of relative atomic masses could be determined. One of the chemists involved in contributing to this list was a Swede named Jöns Jakob Berzelius (1779–1848), who discovered the elements cerium, selenium, silicon, and thorium and developed the modern symbols for the elements used in writing the formulas of compounds.

The Italian chemist Stanislao Cannizzaro (1826–1910) cleared up the confusion in 1860 by doing a series of molar mass determinations that convinced the scientific community that the correct atomic mass of carbon is 12. For more information, see *From Caveman to Chemist* by Hugh Salzberg (American Chemical Society, 1991), p. 223.

2.4 Early Experiments to Characterize the Atom

On the basis of the work of Dalton, Gay-Lussac, Avogadro, and others, chemistry was beginning to make sense. The concept of atoms was clearly a good idea. Inevitably, scientists began to wonder about the nature of the atom. What is an atom made of, and how do the atoms of the various elements differ?

The Electron

The first important experiments that led to an understanding of the composition of the atom were done by the English physicist J. J. Thomson (Fig. 2.6), who studied electrical discharges in partially evacuated tubes called **cathode-ray tubes** (Fig. 2.7) during the period from 1898 to 1903. Thomson found that when high voltage was applied to the tube, a "ray" he called a *cathode ray* (because it emanated from the negative electrode, or cathode) was produced. Because this ray was produced at the negative electrode and was repelled by the negative pole of an applied electric field (see Fig. 2.8), Thomson postulated that the ray was a stream of negatively charged particles, now called **electrons.** From experiments in which he measured the deflection of the beam of electrons in a magnetic field, Thomson determined the *charge-to-mass ratio* of an electron:

$$\frac{e}{m} = -1.76 \times 10^8 \text{ C/g}$$

where e represents the charge on the electron in coulombs (C) and m represents the electron mass in grams.

CHEMICAL IMPACT

Berzelius, Selenium, and Silicon

Jöns Jakob Berzelius was probably the best experimental chemist of his generation and, given the crudeness of his laboratory equipment, maybe the best of all time. Unlike Lavoisier, who could afford to buy the best laboratory equipment available, Berzelius worked with minimal equipment in very plain surroundings. One of Berzelius's students described the Swedish chemist's workplace: "The laboratory consisted of two ordinary rooms with the very simplest arrangements; there were neither furnaces nor hoods, neither water system nor gas. Against the walls stood some closets with the chemicals, in the middle the mercury trough and the blast lamp table. Beside this was the sink consisting of a stone water holder with a stopcock and a pot standing under it. [Next door in the kitchen] stood a small heating furnace."

Comparison of Several of Berzelius's Atomic Masses with the Modern Values

	Atomic Mass	
Element	**Berzelius's Value**	**Current Value**
Chlorine	35.41	35.45
Copper	63.00	63.55
Hydrogen	1.00	1.01
Lead	207.12	207.2
Nitrogen	14.05	14.01
Oxygen	16.00	16.00
Potassium	39.19	39.10
Silver	108.12	107.87
Sulfur	32.18	32.07

In these simple facilities Berzelius performed more than 2000 experiments over a 10-year period to determine accurate atomic masses for the 50 elements then known. His success can be seen from the data in the table at left. These remarkably accurate values attest to his experimental skills and patience.

Besides his table of atomic masses, Berzelius made many other major contributions to chemistry. The most important of these was the invention of a simple set of symbols for the elements along with a system for writing the formulas of compounds to replace the awkward symbolic representations of the alchemists. Although some chemists, including Dalton, objected to the new system, it was gradually adopted and forms the basis of the system we use today.

In addition to these accomplishments, Berzelius discovered the elements cerium, thorium, selenium, and silicon. Of these elements, selenium and silicon are particularly important in today's world. Berzelius discovered selenium in 1817 in connection with his studies of sulfuric acid. For years selenium's toxicity has been known, but only recently have we become aware that it may have a positive effect on human

Visualization: Cathode-Ray Tube

One of Thomson's primary goals in his cathode-ray tube experiments was to gain an understanding of the structure of the atom. He reasoned that since electrons could be produced from electrodes made of various types of metals, *all* atoms must contain electrons. Since atoms were known to be electrically neutral, Thomson further assumed that atoms also must contain some positive charge. Thomson postulated that an atom consisted of a

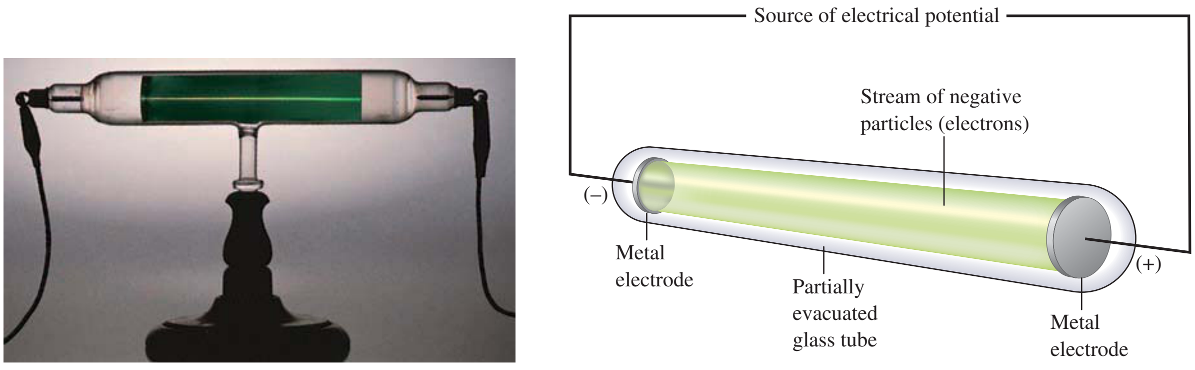
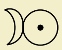
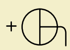


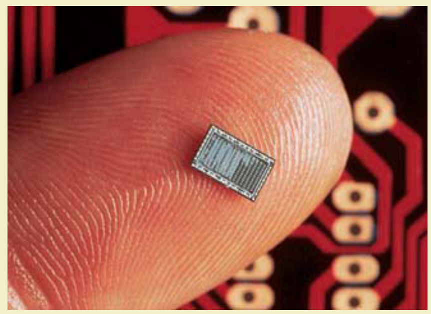
FIGURE 2.7
A cathode-ray tube. The fast-moving electrons excite the gas in the tube, causing a glow between the electrodes. The green color in the photo is due to the response of the screen (coated with zinc sulfide) to the electron beam.

The Alchemists' Symbols for Some Common Elements and Compounds

Substance	Alchemists' Symbol
Silver	☽
Lead	♄
Tin	♃
Platinum	
Sulfuric acid	
Alcohol	
Sea salt	

health. Studies have shown that trace amounts of selenium in the diet may protect people from heart disease and cancer. One study based on data from 27 countries showed an inverse relationship between the cancer death rate and the selenium content of soil in a particular region (low cancer death rate in areas with high selenium content). Another research paper reported an inverse relationship between the selenium content of the blood and the incidence of breast cancer in women. A study reported in 1998 used the toenail clippings of 33,737 men to show that selenium seems to protect against prostate cancer. Selenium is also found in the heart muscle and may play an important role in proper heart function. Because of these and other studies, selenium's reputation has improved, and many scientists are now studying its function in the human body.

Silicon is the second most abundant element in the earth's crust, exceeded only by oxygen. As we will see in Chapter 10, compounds involving silicon bonded to oxygen make up most of the earth's sand, rock, and soil. Berzelius prepared silicon in its pure form in 1824 by heating silicon tetrafluoride (SiF_4) with potassium metal. Today, silicon forms the basis for the modern microelectronics industry centered near San Francisco in a place that has come to be known as "Silicon Valley." The technology of the silicon chip (see figure) with its printed circuits has transformed computers from room-sized monsters with thousands of unreliable vacuum tubes to desktop and notebook-sized units with trouble-free "solid-state" circuitry.



A silicon chip.

See E. J. Holmyard, *Alchemy* (New York: Penguin Books, 1968).

diffuse cloud of positive charge with the negative electrons embedded randomly in it. This model, shown in Fig. 2.9, is often called the *plum pudding model* because the electrons are like raisins dispersed in a pudding (the positive charge cloud), as in plum pudding, a favorite English dessert.

In 1909 Robert Millikan (1868–1953), working at the University of Chicago, performed very clever experiments involving charged oil drops. These experiments allowed

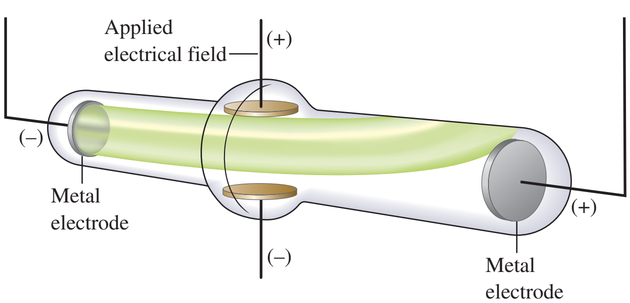


FIGURE 2.8
Deflection of cathode rays by an applied electric field.

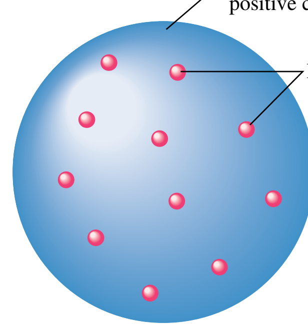


FIGURE 2.9
The plum pudding model of the atom.

Visualization: Millikan's Oil Drop Experiment

A technician using a scanner to monitor the uptake of radioactive iodine in a patient's thyroid.

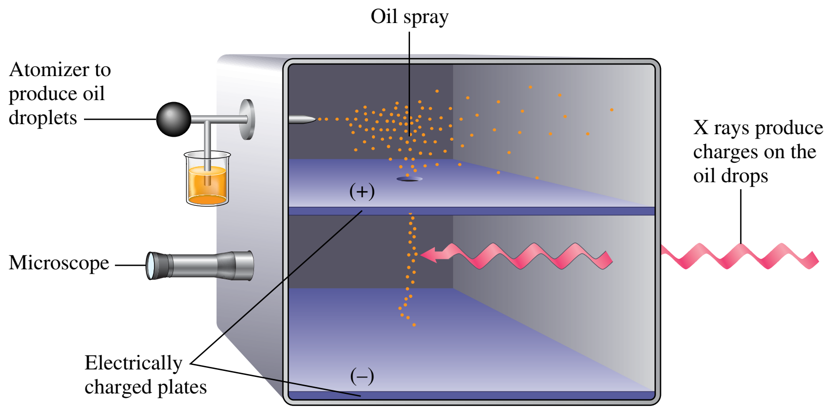


FIGURE 2.10
A schematic representation of the apparatus Millikan used to determine the charge on the electron. The fall of charged oil droplets due to gravity can be halted by adjusting the voltage across the two plates. This voltage and the mass of the oil drop can then be used to calculate the charge on the oil drop. Millikan's experiments showed that the charge on an oil drop is always a whole-number multiple of the electron charge.

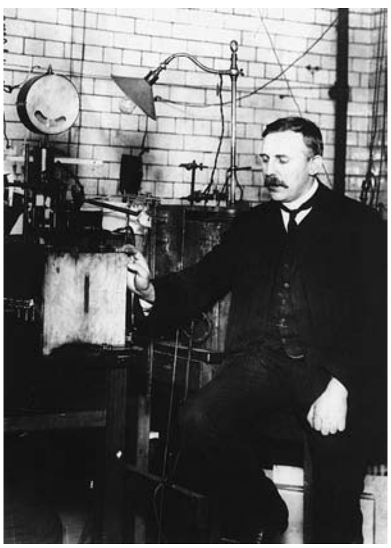


FIGURE 2.11
Ernest Rutherford (1871–1937) was born on a farm in New Zealand. In 1895 he placed second in a scholarship competition to attend Cambridge University but was awarded the scholarship when the winner decided to stay home and get married. As a scientist in England, Rutherford did much of the early work on characterizing radioactivity. He named the α and β particles and the γ ray and coined the term *half-life* to describe an important attribute of radioactive elements. His experiments on the behavior of α particles striking thin metal foils led him to postulate the nuclear atom. He also invented the name *proton* for the nucleus of the hydrogen atom. He received the Nobel Prize in chemistry in 1908.

him to determine the magnitude of the electron charge (see Fig. 2.10). With this value and the charge-to-mass ratio determined by Thomson, Millikan was able to calculate the mass of the electron as 9.11×10^{-31} kilogram.

Radioactivity

In the late nineteenth century scientists discovered that certain elements produce high-energy radiation. For example, in 1896 the French scientist Henri Becquerel found accidentally that a piece of a mineral containing uranium could produce its image on a photographic plate in the absence of light. He attributed this phenomenon to a spontaneous emission of radiation by the uranium, which he called **radioactivity.** Studies in the early twentieth century demonstrated three types of radioactive emission: gamma (γ) rays, beta (β) particles, and alpha (α) particles. A γ ray is high-energy "light"; a β particle is a high-speed electron; and an α particle has a 2+ charge, that is, a charge twice that of the electron and with the opposite sign. The mass of an α particle is 7300 times that of the electron. More modes of radioactivity are now known, and we will discuss them in Chapter 18. Here we will consider only α particles because they were used in some crucial early experiments.

The Nuclear Atom

In 1911 Ernest Rutherford (Fig. 2.11), who performed many of the pioneering experiments to explore radioactivity, carried out an experiment to test Thomson's plum pudding model. The experiment involved directing α particles at a thin sheet of metal foil, as illustrated in Fig. 2.12. Rutherford reasoned that if Thomson's model were accurate, the massive α particles should crash through the thin foil like cannonballs through gauze, as shown in Fig. 2.13(a). He expected the α particles to travel through the foil with, at the most, very minor deflections in their paths. The results of the experiment were very different from those Rutherford anticipated. Although most of the α particles passed straight through, many of the particles were deflected at large angles, as shown in Fig. 2.13(b), and some were reflected, never hitting the detector. This outcome was a great surprise to Rutherford. (He wrote that this result was comparable with shooting a howitzer at a piece of paper and having the shell reflected back.)

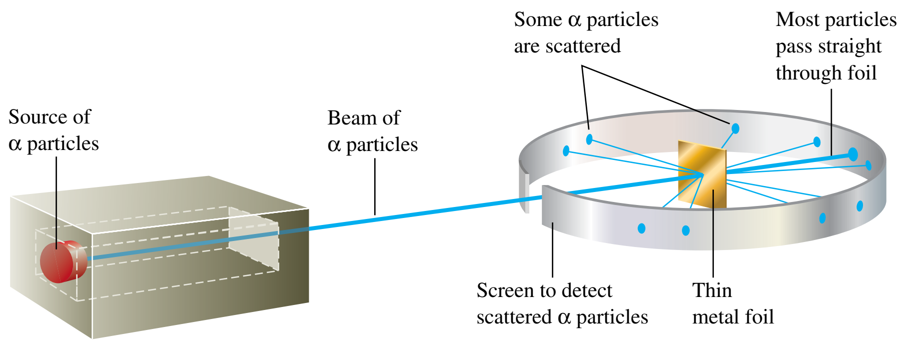


FIGURE 2.12
Rutherford's experiment on α-particle bombardment of metal foil.

Rutherford knew from these results that the plum pudding model for the atom could not be correct. The large deflections of the α particles could be caused only by a center of concentrated positive charge that contains most of the atom's mass, as illustrated in Fig. 2.13(b). Most of the α particles pass directly through the foil because the atom is mostly open space. The deflected α particles are those that had a "close encounter" with the massive positive center of the atom, and the few reflected α particles are those that made a "direct hit" on the much more massive positive center.

In Rutherford's mind these results could be explained only in terms of a **nuclear atom**—an atom with a dense center of positive charge (the **nucleus**) with electrons moving around the nucleus at a distance that is large relative to the nuclear radius.

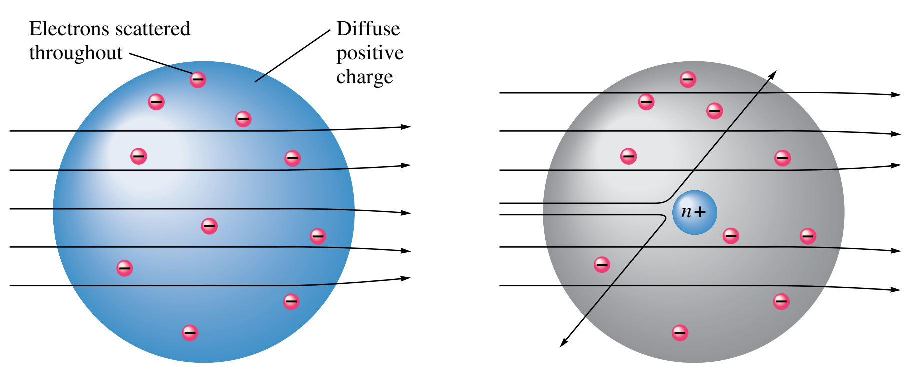
2.5 The Modern View of Atomic Structure: An Introduction

The forces that bind the positively charged protons in the nucleus will be discussed in Chapter 18.

In the years since Thomson and Rutherford, a great deal has been learned about atomic structure. Because much of this material will be covered in detail in later chapters, only an introduction will be given here. The simplest view of the atom is that it consists of a tiny nucleus (with a diameter of about 10^{-13} cm) and electrons that move about the nucleus at an average distance of about 10^{-8} cm from it (see Fig. 2.14).

As we will see later, the chemistry of an atom mainly results from its electrons. For this reason, chemists can be satisfied with a relatively crude nuclear model. The nucleus is assumed to contain **protons,** which have a positive charge equal in magnitude to the electron's negative charge, and **neutrons,** which have virtually the same mass as a proton but no charge. The masses and charges of the electron, proton, and neutron are shown in Table 2.1.

WWW Visualization: Gold Foil Experiment



(a) (b)

FIGURE 2.13
(a) The expected results of the metal foil experiment if Thomson's model were correct. (b) Actual results.

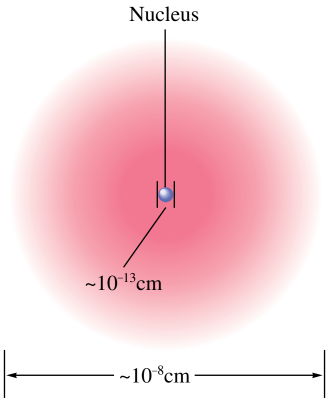


FIGURE 2.14
A nuclear atom viewed in cross section. Note that this drawing is not to scale.

TABLE 2.1 The Mass and Charge of the Electron, Proton, and Neutron		
Particle	**Mass**	**Charge***
Electron	9.11×10^{-31} kg	1−
Proton	1.67×10^{-27} kg	1+
Neutron	1.67×10^{-27} kg	None

*The magnitude of the charge of the electron and the proton is 1.60×10^{-19} C.

Two striking things about the nucleus are its small size compared with the overall size of the atom and its extremely high density. The tiny nucleus accounts for almost all the atom's mass. Its great density is dramatically demonstrated by the fact that a piece of nuclear material about the size of a pea would have a mass of 250 million tons!

The *chemistry* of an atom arises from its electrons.

An important question to consider at this point is, *"If all atoms are composed of these same components, why do different atoms have different chemical properties?"* The answer to this question lies in the number and the arrangement of the electrons. The electrons constitute most of the atomic volume and thus are the parts that "intermingle" when atoms combine to form molecules. Therefore, the number of electrons possessed by a given atom greatly affects its ability to interact with other atoms. As a result, the atoms of different elements, which have different numbers of protons and electrons, show different chemical behavior.

If the atomic nucleus were the size of this ball bearing, a typical atom would be the size of this stadium.

A sodium atom has 11 protons in its nucleus. Since atoms have no net charge, the number of electrons must equal the number of protons. Therefore, a sodium atom has 11 electrons moving around its nucleus. It is *always* true that a sodium atom has 11 protons and 11 electrons. However, each sodium atom also has neutrons in its nucleus, and different types of sodium atoms exist that have different numbers of neutrons. For example, consider the sodium atoms represented in Fig. 2.15. These two atoms are **isotopes,** or *atoms with the same number of protons but different numbers of neutrons.* Note that the symbol for one particular type of sodium atom is written

Mass number ↘
$^{23}_{11}\mathrm{Na}$ ← Element symbol
Atomic number ↗

where the **atomic number** Z (number of protons) is written as a subscript, and the **mass number** A (the total number of protons and neutrons) is written as a superscript. (The particular atom represented here is called "sodium twenty-three." It has 11 electrons, 11 protons, and 12 neutrons.) Because the chemistry of an atom is due to its electrons, isotopes show almost identical chemical properties. In nature most elements contain mixtures of isotopes.

Mass number ↘
Atomic number → $^{A}_{Z}\mathrm{X}$ ← Element symbol

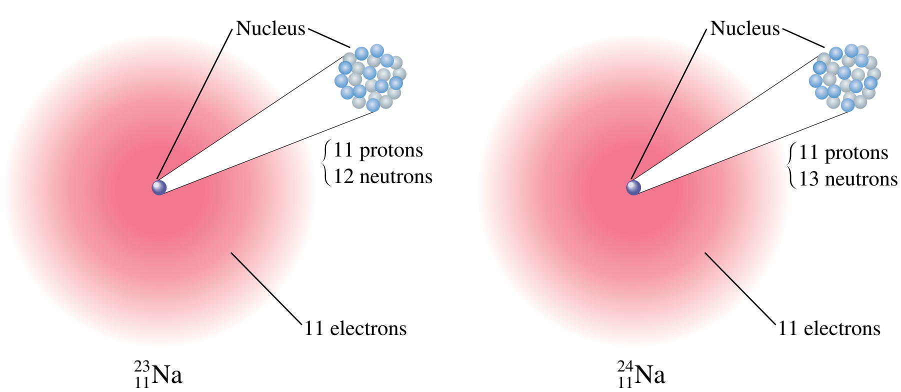


FIGURE 2.15
Two isotopes of sodium. Both have 11 protons and 11 electrons, but they differ in the number of neutrons in their nuclei.

CHEMICAL IMPACT

Reading the History of Bogs

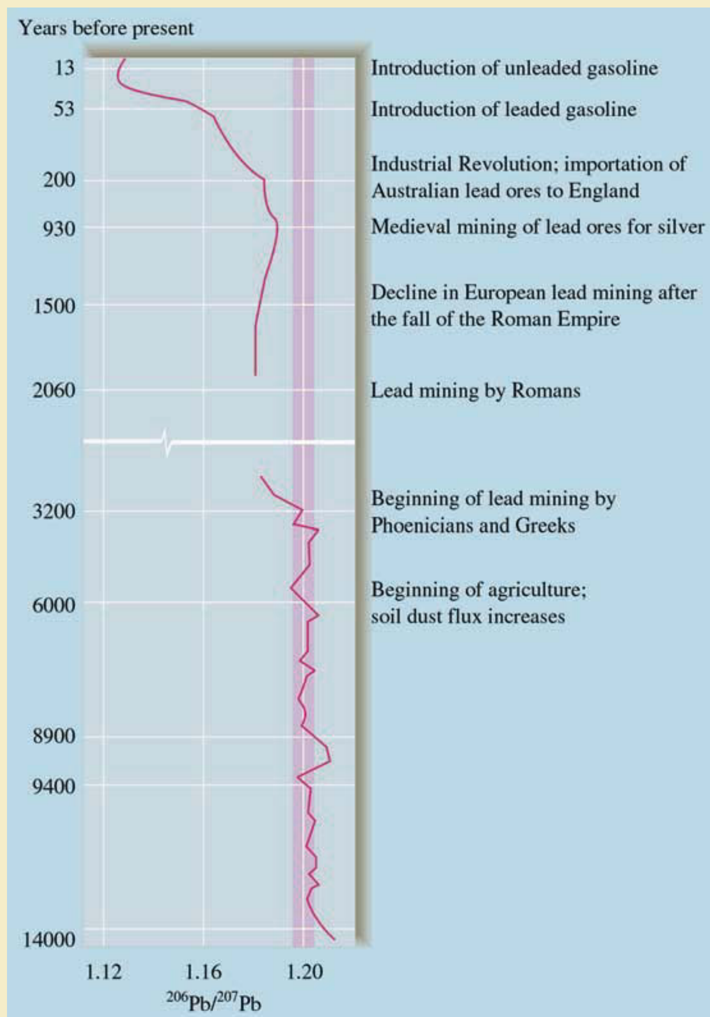
Scientists often "read" the history of the earth and its inhabitants using very different "books" than traditional historians. For example, the disappearance of the dinosaurs 65 million years ago in an "instant" of geological time was a great mystery until unusually high iridium and osmium levels were discovered at a position in the earth's crust corresponding to that time. These high levels of iridium and osmium suggested that an extraterrestrial object had struck the earth 65 million years ago with catastrophic results for the dinosaurs. Since then, the huge buried crater caused by the object has been discovered on the Yucatan Peninsula, and virtually everyone is now convinced that this is the correct explanation for the disappearing dinosaurs.

History is also being "read" by scientists studying ice cores from glaciers in Iceland. Now Swiss scientists have found that ancient peat bogs can furnish a reliable historical record. Geochemist William Shotyk of the University of Bern has found a 15,000-year window on history by analyzing the lead content of core samples from a Swiss mountainside peat bog [*Science* 281 (1998): 1635]. Various parts of the core samples were dated by ^{14}C dating techniques (see Chapter 18, Section 18.4, for more information) and analyzed for their scandium and lead contents. Also, the $^{206}Pb/^{207}Pb$ ratio was measured for each sample. These data are represented in the accompanying figure. Notice that the $^{206}Pb/^{207}Pb$ ratio remains very close to 1.20 (see the red band in the figure) from 14,000 years to 3200 years. The value of 1.20 is the same as the average $^{206}Pb/^{207}Pb$ ratio in the earth's soil.

The core also reveals that the total lead and scandium levels increased simultaneously at the 6000-year mark but that the $^{206}Pb/^{207}Pb$ ratio remained close to 1.20. This coincides with the beginning of agriculture in Europe, which caused more soil dust to enter the atmosphere.

Significantly, about 3000 years ago the $^{206}Pb/^{207}Pb$ ratio decreased markedly. This also corresponds in the core sample to an increase in total lead content out of proportion to the increase in scandium. This indicates the lead no longer resulted from soil dust but from other activities of humans—lead mining had begun. Since the 3000-year mark, the $^{206}Pb/^{207}Pb$ ratio has remained well below 1.20, indicating that human use of lead ores has become the dominant source of airborne lead. This is confirmed by the sharp decline in the ratio beginning 200 years ago that corresponds to the importation into England of Australian lead ores having low $^{206}Pb/^{207}Pb$ ratios.

So far only lead has been used to read the history in the bog. However, Shotyk's group is also measuring the changes in the levels of copper, zinc, cadmium, arsenic, mercury, and antimony. More interesting stories are sure to follow.



Geochemist William Shotyk's analysis of the lead content of ice core samples reveals a 15,000-year history of lead levels. (Note: Dates are based on calibrated radiocarbon dating. Because the core was retrieved in two segments, a break in data occurs between 2060 and 3200 years before present.)

Sample Exercise 2.2 Writing the Symbols for Atoms

Write the symbol for the atom that has an atomic number of 9 and a mass number of 19. How many electrons and how many neutrons does this atom have?

Solution

The atomic number 9 means the atom has 9 protons. This element is called *fluorine,* symbolized by F. The atom is represented as

$$^{19}_{9}\mathrm{F}$$

and is called "fluorine nineteen." Since the atom has 9 protons, it also must have 9 electrons to achieve electrical neutrality. The mass number gives the total number of protons and neutrons, which means that this atom has 10 neutrons.

See Exercises 2.43 through 2.46.

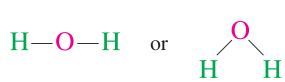
2.6 Molecules and Ions

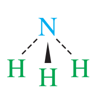
From a chemist's viewpoint, the most interesting characteristic of an atom is its ability to combine with other atoms to form compounds. It was John Dalton who first recognized that chemical compounds are collections of atoms, but he could not determine the structure of atoms or their means for binding to each other. During the twentieth century we learned that atoms have electrons and that these electrons participate in bonding one atom to another. We will discuss bonding thoroughly in Chapters 8 and 9; here we will introduce some simple bonding ideas that will be useful in the next few chapters.

The forces that hold atoms together in compounds are called **chemical bonds.** One way that atoms can form bonds is by *sharing electrons.* These bonds are called **covalent bonds,** and the resulting collection of atoms is called a **molecule.** Molecules can be represented in several different ways. The simplest method is the **chemical formula,** in which the symbols for the elements are used to indicate the types of atoms present and subscripts are used to indicate the relative numbers of atoms. For example, the formula for carbon dioxide is CO_2, meaning that each molecule contains 1 atom of carbon and 2 atoms of oxygen.

Visualization: Covalent Bonding

Examples of molecules that contain covalent bonds are hydrogen (H_2), water (H_2O), oxygen (O_2), ammonia (NH_3), and methane (CH_4). More information about a molecule is given by its **structural formula,** in which the individual bonds are shown (indicated by lines). Structural formulas may or may not indicate the actual shape of the molecule. For example, water might be represented as



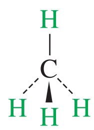


Ammonia

The structure on the right shows the actual shape of the water molecule. Scientists know from experimental evidence that the molecule looks like this. (We will study the shapes of molecules further in Chapter 8.) The structural formula for ammonia is shown in the margin at left.

Note that atoms connected to the central atom by dashed lines are behind the plane of the paper, and atoms connected to the central atom by wedges are in front of the plane of the paper.

In a compound composed of molecules, the individual molecules move around as independent units. For example, a molecule of methane gas can be represented in several ways. The structural formula for methane (CH_4) is shown in Fig. 2.16. The **space-filling**



Methane

FIGURE 2.16
The structural formula for methane.

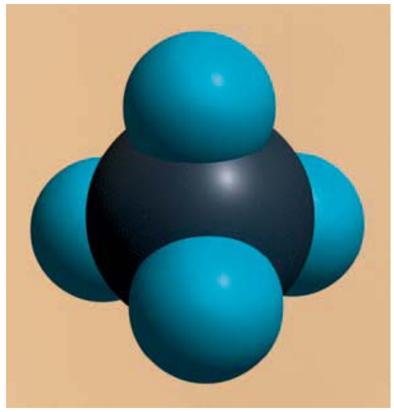


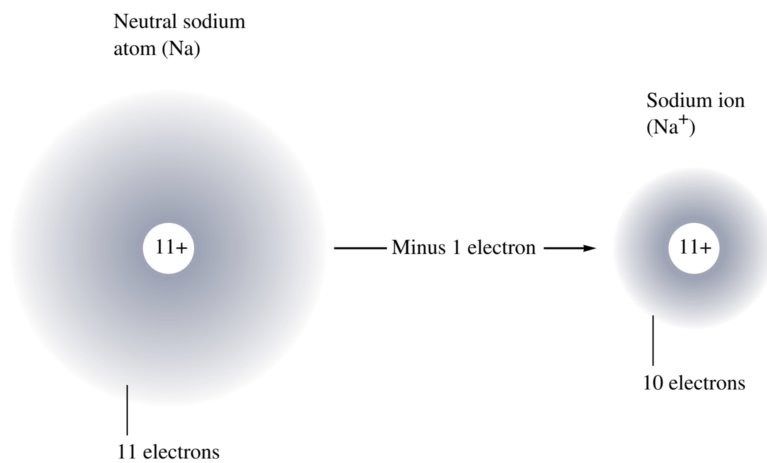
FIGURE 2.17
Space-filling model of methane. This type of model shows both the relative sizes of the atoms in the molecule and their spatial relationships.

FIGURE 2.18
Ball-and-stick model of methane.

model of methane, which shows the relative sizes of the atoms as well as their relative orientation in the molecule, is given in Fig. 2.17. **Ball-and-stick models** are also used to represent molecules. The ball-and-stick structure of methane is shown in Fig. 2.18.

A second type of chemical bond results from attractions among ions. An **ion** is an atom or group of atoms that has a net positive or negative charge. The best-known ionic compound is common table salt, or sodium chloride, which forms when neutral chlorine and sodium react.

To see how the ions are formed, consider what happens when an electron is transferred from a sodium atom to a chlorine atom (the neutrons in the nuclei will be ignored):

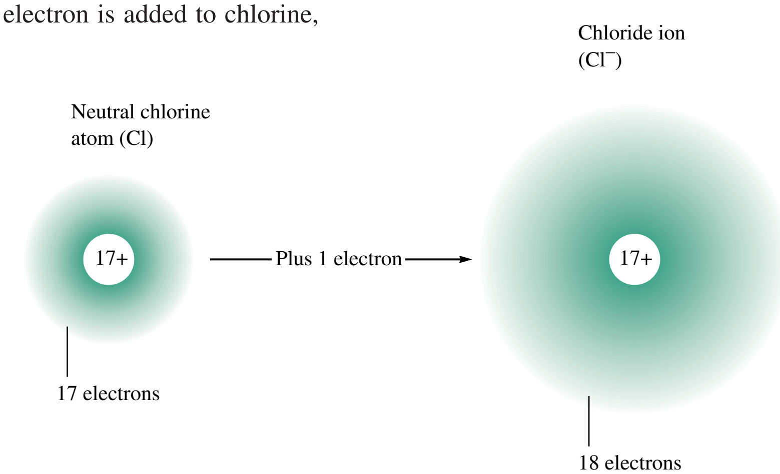


Na^+ is usually called the *sodium ion* rather than the sodium cation. Also Cl^- is called the *chloride ion* rather than the chloride anion. In general, when a specific ion is referred to, the word *ion* rather than cation or anion is used.

With one electron stripped off, the sodium, with its 11 protons and only 10 electrons, now has a net 1+ charge—it has become a *positive ion.* A positive ion is called a **cation.** The sodium ion is written as Na^+, and the process can be represented in shorthand form as

$$Na \longrightarrow Na^+ + e^-$$

If an electron is added to chlorine,



the 18 electrons produce a net 1− charge; the chlorine has become an *ion with a negative charge*—an **anion.** The chloride ion is written as Cl^-, and the process is represented as

$$Cl + e^- \longrightarrow Cl^-$$

Because anions and cations have opposite charges, they attract each other. This *force of attraction between oppositely charged ions* is called **ionic bonding.** As illustrated in Fig. 2.19, sodium metal and chlorine gas (a green gas composed of Cl_2 molecules) react

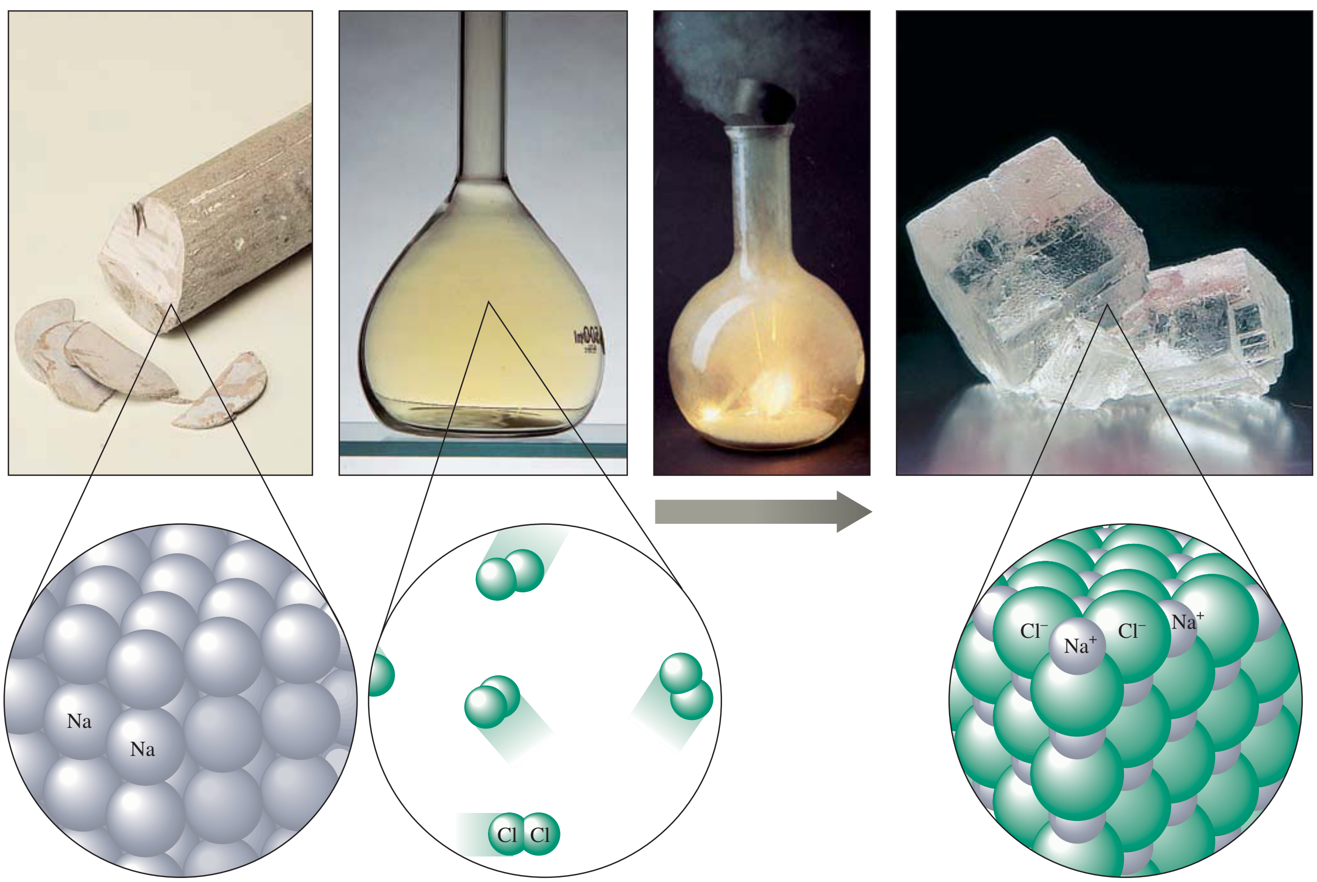


FIGURE 2.19
Sodium metal (which is so soft it can be cut with a knife and which consists of individual sodium atoms) reacts with chlorine gas (which contains Cl_2 molecules) to form solid sodium chloride (which contains Na^+ and Cl^- ions packed together).

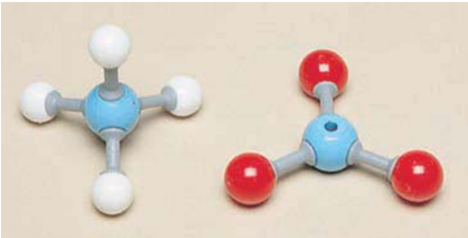


FIGURE 2.20
Ball-and-stick models of the ammonium ion (NH_4^+) and the nitrate ion (NO_3^-).

to form solid sodium chloride, which contains many Na^+ and Cl^- ions packed together and forms the beautiful colorless cubic crystals shown in Fig. 2.19.

A solid consisting of oppositely charged ions is called an **ionic solid,** or a **salt.** Ionic solids can consist of simple ions, as in sodium chloride, or of **polyatomic** (many atom) **ions,** as in ammonium nitrate (NH_4NO_3), which contains ammonium ions (NH_4^+) and nitrate ions (NO_3^-). The ball-and-stick models of these ions are shown in Fig. 2.20.

2.7 An Introduction to the Periodic Table

In a room where chemistry is taught or practiced, a chart called the **periodic table** is almost certain to be found hanging on the wall. This chart shows all the known elements and gives a good deal of information about each. As our study of chemistry progresses, the usefulness of the periodic table will become more obvious. This section will simply introduce it to you.

Visualization: Comparison of a Molecular Compound and an Ionic Compound

A simplified version of the periodic table is shown in Fig. 2.21. The letters in the boxes are the symbols for the elements; these abbreviations are based on the current element names or the original names (see Table 2.2). The number shown above each symbol is the *atomic number* (number of protons) for that element. For example, carbon (C) has atomic number 6, and lead (Pb) has atomic number 82. Most of the elements are **metals.** Metals have characteristic physical properties such as efficient conduction of heat and electricity, malleability (they can be hammered into thin sheets), ductility (they can be pulled into wires), and (often) a lustrous appearance. Chemically, metals tend to *lose* electrons to form positive ions. For example, copper is a typical metal. It is lustrous (although it tarnishes readily); it is an excellent conductor of electricity (it is widely used in electrical wires); and it is readily formed into various shapes, such as pipes for water systems. Copper is also found in many salts, such as the beautiful blue copper sulfate, in which copper is present as Cu^{2+} ions. Copper is a member of the transition metals—the metals shown in the center of the periodic table.

Metals tend to form positive ions; nonmetals tend to form negative ions.

The relatively few **nonmetals** appear in the upper-right corner of the table (to the right of the heavy line in Fig. 2.21), except hydrogen, a nonmetal that resides in the upper-left corner. The nonmetals lack the physical properties that characterize the metals. Chemically, they tend to *gain* electrons in reactions with metals to form negative ions. Nonmetals often bond to each other by forming covalent bonds. For example, chlorine is a typical nonmetal. Under normal conditions it exists as Cl_2 molecules; it reacts with metals to form salts containing Cl^- ions (NaCl, for example); and it forms covalent bonds with nonmetals (for example, hydrogen chloride gas, HCl).

Elements in the same vertical column in the periodic table form a *group* (or *family*) and generally have similar properties.

The periodic table is arranged so that elements in the same vertical columns (called **groups** or **families**) have *similar chemical properties.* For example, all of the **alkali metals,** members of Group 1A—lithium (Li), sodium (Na), potassium (K), rubidium (Rb), cesium (Cs), and francium (Fr)—are very active elements that readily form ions with a 1+ charge when they react with nonmetals. The members of Group 2A—beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra)—are called the **alkaline earth metals.** They all form ions with a 2+ charge when they react with nonmetals. The **halogens,** the members of Group 7A—fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At)—all form diatomic molecules. Fluorine, chlorine, bromine, and iodine all react with metals to form salts containing ions with a 1− charge (F^-, Cl^-, Br^-, and I^-). The members of Group 8A—helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn)—are known as the **noble gases.** They all exist under normal conditions as monatomic (single-atom) gases and have little chemical reactivity.

Samples of chlorine gas, liquid bromine, and solid iodine.

Alkaline earth metals

Noble gases

Halogens

Transition metals

Alkali metals

1 (1A)	2 (2A)	3	4	5	6	7	8	9	10	11	12	13 (3A)	14 (4A)	15 (5A)	16 (6A)	17 (7A)	18 (8A)
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La*	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac†	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Uub	113 Uut	114 Uuq	115 Uup			

*Lanthanides	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
†Actinides	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

FIGURE 2.21
The periodic table.

TABLE 2.2 The Symbols for the Elements That Are Based on the Original Names

Current Name	Original Name	Symbol
Antimony	Stibium	Sb
Copper	Cuprum	Cu
Iron	Ferrum	Fe
Lead	Plumbum	Pb
Mercury	Hydrargyrum	Hg
Potassium	Kalium	K
Silver	Argentum	Ag
Sodium	Natrium	Na
Tin	Stannum	Sn
Tungsten	Wolfram	W

CHEMICAL IMPACT

Hassium Fits Right in

Hassium, element 108, does not exist in nature but must be made in a particle accelerator. It was first created in 1984 and can be made by shooting magnesium-26 ($^{26}_{12}Mg$) atoms at curium-248 ($^{248}_{96}Cm$) atoms. The collisions between these atoms produce some hassium-265 ($^{265}_{108}Hs$) atoms. The position of hassium in the periodic table (see Fig. 2.21) in the vertical column containing iron, ruthenium, and osmium suggests that hassium should have chemical properties similar to these metals. However, it is not easy to test this prediction—only a few atoms of hassium can be made at a given time and they last for only about 9 seconds. Imagine having to get your next lab experiment done in 9 seconds!

Amazingly, a team of chemists from the Lawrence Berkeley National Laboratory in California, the Paul Scherrer Institute and the University of Bern in Switzerland, and the Institute of Nuclear Chemistry in Germany have done experiments to characterize the chemical behavior of hassium. For example, they have observed that hassium atoms react with oxygen to form a hassium oxide compound of the type expected from its position on the periodic table. The team has also measured other properties of hassium, including the energy released as it undergoes nuclear decay to another atom.

This work would have surely pleased Dmitri Mendeleev (see Fig. 7.23), who originally developed the periodic table and showed its power to predict chemical properties.

Note from Fig. 2.21 that alternate sets of symbols are used to denote the groups. The symbols 1A through 8A are the traditional designations, whereas the numbers 1 to 18 have been suggested recently. In this text the 1A to 8A designations will be used.

The horizontal rows of elements in the periodic table are called **periods.** Horizontal row 1 is called the *first period* (it contains H and He); row 2 is called the *second period* (elements Li through Ne); and so on.

We will learn much more about the periodic table as we continue with our study of chemistry. Meanwhile, when an element is introduced in this text, you should always note its position on the periodic table.

Another format of the periodic table will be discussed in Section 7.11.

2.8 Naming Simple Compounds

When chemistry was an infant science, there was no system for naming compounds. Names such as sugar of lead, blue vitrol, quicklime, Epsom salts, milk of magnesia, gypsum, and laughing gas were coined by early chemists. Such names are called *common names.* As chemistry grew, it became clear that using common names for compounds would lead to unacceptable chaos. Nearly 5 million chemical compounds are currently known. Memorizing common names for these compounds would be an impossible task.

The solution, of course, is to adopt a *system* for naming compounds in which the name tells something about the composition of the compound. After learning the system, a chemist given a formula should be able to name the compound or, given a name, should be able to construct the compound's formula. In this section we will specify the most important rules for naming compounds other than organic compounds (those based on chains of carbon atoms).

We will begin with the systems for naming inorganic **binary compounds**—compounds composed of two elements—which we classify into various types for easier recognition. We will consider both ionic and covalent compounds.

TABLE 2.3 Common Monatomic Cations and Anions

Cation	Name	Anion	Name
H^+	Hydrogen	H^-	Hydride
Li^+	Lithium	F^-	Fluoride
Na^+	Sodium	Cl^-	Chloride
K^+	Potassium	Br^-	Bromide
Cs^+	Cesium	I^-	Iodide
Be^{2+}	Beryllium	O^{2-}	Oxide
Mg^{2+}	Magnesium	S^{2-}	Sulfide
Ca^{2+}	Calcium	N^{3-}	Nitride
Ba^{2+}	Barium	P^{3-}	Phosphide
Al^{3+}	Aluminum		
Ag^+	Silver		

Binary Ionic Compounds (Type I)

Binary ionic compounds contain a positive ion (cation) always written first in the formula and a negative ion (anion). In naming these compounds, the following rules apply:

1. The cation is always named first and the anion second.
2. A monatomic (meaning "one-atom") cation takes its name from the name of the element. For example, Na^+ is called sodium in the names of compounds containing this ion.
3. A monatomic anion is named by taking the root of the element name and adding *-ide.* Thus the Cl^- ion is called chloride.

A monatomic cation has the same name as its parent element.

Some common monatomic cations and anions and their names are given in Table 2.3.

The rules for naming binary ionic compounds are illustrated by the following examples:

In formulas of ionic compounds, simple ions are represented by the element symbol: Cl means Cl^-, Na means Na^+, and so on.

Compound	Ions Present	Name
NaCl	Na^+, Cl^-	Sodium chloride
KI	K^+, I^-	Potassium iodide
CaS	Ca^{2+}, S^{2-}	Calcium sulfide
Li_3N	Li^+, N^{3-}	Lithium nitride
CsBr	Cs^+, Br^-	Cesium bromide
MgO	Mg^{2+}, O^{2-}	Magnesium oxide

Sample Exercise 2.3

Naming Type I Binary Compounds

Name each binary compound.

a. CsF **b.** $AlCl_3$ **c.** LiH

Solution

a. CsF is cesium fluoride.
b. $AlCl_3$ is aluminum chloride.
c. LiH is lithium hydride.

Notice that, in each case, the cation is named first, and then the anion is named.

See Exercise 2.55.

Visualization: Formation of Ionic Compounds

TABLE 2.4 Common Type II Cations

Ion	Systematic Name
Fe^{3+}	Iron(III)
Fe^{2+}	Iron(II)
Cu^{2+}	Copper(II)
Cu^{+}	Copper(I)
Co^{3+}	Cobalt(III)
Co^{2+}	Cobalt(II)
Sn^{4+}	Tin(IV)
Sn^{2+}	Tin(II)
Pb^{4+}	Lead(IV)
Pb^{2+}	Lead(II)
Hg^{2+}	Mercury(II)
Hg_2^{2+}*	Mercury(I)
Ag^{+}	Silver†
Zn^{2+}	Zinc†
Cd^{2+}	Cadmium†

*Note that mercury(I) ions always occur bound together to form Hg_2^{2+} ions.
†Although these are transition metals, they form only one type of ion, and a Roman numeral is not used.

Formulas from Names

So far we have started with the chemical formula of a compound and decided on its systematic name. The reverse process is also important. For example, given the name calcium hydroxide, we can write the formula as $Ca(OH)_2$ because we know that calcium forms only Ca^{2+} ions and that, since hydroxide is OH^-, two of these anions will be required to give a neutral compound.

Binary Ionic Compounds (Type II)

In the binary ionic compounds considered earlier (Type I), the metal present forms only a single type of cation. That is, sodium forms only Na^+, calcium forms only Ca^{2+}, and so on. However, as we will see in more detail later in the text, there are many metals that form more than one type of positive ion and thus form more than one type of ionic compound with a given anion. For example, the compound $FeCl_2$ contains Fe^{2+} ions, and the compound $FeCl_3$ contains Fe^{3+} ions. In a case such as this, the *charge on the metal ion must be specified.* The systematic names for these two iron compounds are iron(II) chloride and iron(III) chloride, respectively, where the *Roman numeral indicates the charge of the cation.*

Another system for naming these ionic compounds that is seen in the older literature was used for metals that form only two ions. *The ion with the higher charge has a name ending in* -ic, *and the one with the lower charge has a name ending in* -ous. In this system, for example, Fe^{3+} is called the ferric ion, and Fe^{2+} is called the ferrous ion. The names for $FeCl_3$ and $FeCl_2$ are then ferric chloride and ferrous chloride, respectively. In this text we will use the system that employs Roman numerals. Table 2.4 lists the systematic names for many common type II cations.

Sample Exercise 2.4

Formulas from Names for Type I Binary Compounds

Given the following systematic names, write the formula for each compound:

a. potassium iodide
b. calcium oxide
c. gallium bromide

Solution

Name	Formula	Comments
a. potassium iodide	KI	Contains K^+ and I^-.
b. calcium oxide	CaO	Contains Ca^{2+} and O^{2-}.
c. gallium bromide	$GaBr_3$	Contains Ga^{3+} and Br^-. Must have $3Br^-$ to balance charge of Ga^{3+}.

See Exercise 2.55.

Sample Exercise 2.5

Naming Type II Binary Compounds

1. Give the systematic name for each of the following compounds:

a. CuCl **b.** HgO **c.** Fe_2O_3

2. Given the following systematic names, write the formula for each compound:

a. Manganese(IV) oxide
b. Lead(II) chloride

Type II binary ionic compounds contain a metal that can form more than one type of cation.

A compound must be electrically neutral.

Solution

All of these compounds include a metal that can form more than one type of cation. Thus we must first determine the charge on each cation. This can be done by recognizing that a compound must be electrically neutral; that is, the positive and negative charges must exactly balance.

1.

Formula	Name	Comments
a. CuCl	Copper(I) chloride	Because the anion is Cl^-, the cation must be Cu^+ (for charge balance), which requires a Roman numeral I.
b. HgO	Mercury(II) oxide	Because the anion is O^{2-}, the cation must be Hg^{2+} [mercury(II)].
c. Fe_2O_3	Iron(III) oxide	The three O^{2-} ions carry a total charge of 6−, so two Fe^{3+} ions [iron(III)] are needed to give a 6+ charge.

2.

Name	Formula	Comments
a. Manganese(IV) oxide	MnO_2	Two O^{2-} ions (total charge 4−) are required by the Mn^{4+} ion [manganese(IV)].
b. Lead(II) chloride	$PbCl_2$	Two Cl^- ions are required by the Pb^{2+} ion [lead(II)] for charge balance.

See Exercise 2.56.

A compound containing a transition metal usually requires a Roman numeral in its name.

Note that the use of a Roman numeral in a systematic name is required only in cases where more than one ionic compound forms between a given pair of elements. This case most commonly occurs for compounds containing transition metals, which often form more than one cation. *Elements that form only one cation do not need to be identified by a Roman numeral.* Common metals that do not require Roman numerals are the Group 1A elements, which form only 1+ ions; the Group 2A elements, which form only 2+ ions; and aluminum, which forms only Al^{3+}. The element silver deserves special mention at this point. In virtually all its compounds silver is found as the Ag^+ ion. Therefore, although silver is a transition metal (and can potentially form ions other than Ag^+), silver compounds are usually named without a Roman numeral. Thus AgCl is typically called silver chloride rather than silver(I) chloride, although the latter name is technically correct. Also, a Roman numeral is not used for zinc compounds, since zinc forms only the Zn^{2+} ion.

As shown in Sample Exercise 2.5, when a metal ion is present that forms more than one type of cation, the charge on the metal ion must be determined by balancing the positive and negative charges of the compound. To do this you must be able to recognize the common cations and anions and know their charges (see Tables 2.3 and 2.5).

Crystals of copper(II) sulfate.

Sample Exercise 2.6 Naming Binary Compounds

1. Give the systematic name for each of the following compounds:

a. $CoBr_2$ **b.** $CaCl_2$ **c.** Al_2O_3

2. Given the following systematic names, write the formula for each compound:

a. Chromium(III) chloride
b. Gallium iodide

Solution

1.

Formula	**Name**	**Comments**
a. $CoBr_2$	Cobalt(II) bromide	Cobalt is a transition metal; the compound name must have a Roman numeral. The two Br^- ions must be balanced by a Co^{2+} ion.
b. $CaCl_2$	Calcium chloride	Calcium, an alkaline earth metal, forms only the Ca^{2+} ion. A Roman numeral is not necessary.
c. Al_2O_3	Aluminum oxide	Aluminum forms only the Al^{3+} ion. A Roman numeral is not necessary.

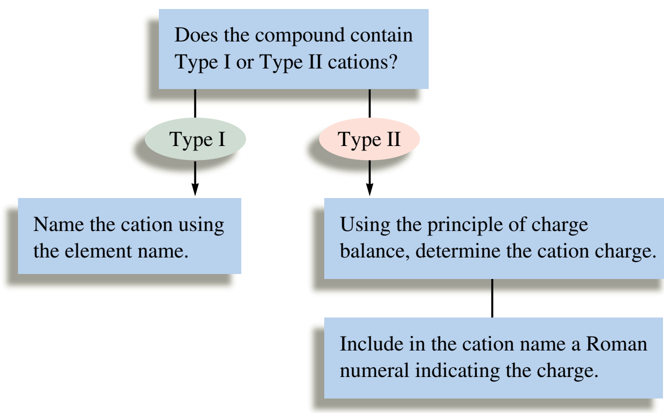
2.

Name	**Formula**	**Comments**
a. Chromium(III) chloride	$CrCl_3$	Chromium(III) indicates that Cr^{3+} is present, so 3 Cl^- ions are needed for charge balance.
b. Gallium iodide	GaI_3	Gallium always forms 3+ ions, so 3 I^- ions are required for charge balance.

See Exercises 2.57 and 2.58.

Various chromium compounds dissolved in water. From left to right: $CrCl_2$, $K_2Cr_2O_7$, $Cr(NO_3)_3$, $CrCl_3$, K_2CrO_4.

The following flowchart is useful when you are naming binary ionic compounds:



The common Type I and Type II ions are summarized in Fig. 2.22. Also shown in Fig. 2.22 are the common monatomic ions.

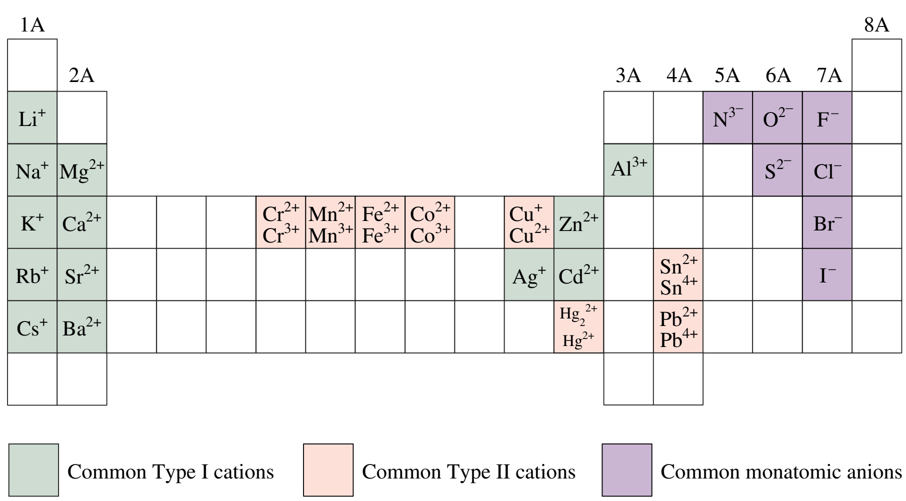


FIGURE 2.22
The common cations and anions.

TABLE 2.5 Common Polyatomic Ions

Ion	Name	Ion	Name
Hg_2^{2+}	Mercury(I)	NCS^-	Thiocyanate
NH_4^+	Ammonium	CO_3^{2-}	Carbonate
NO_2^-	Nitrite	HCO_3^-	Hydrogen carbonate (bicarbonate is a widely used common name)
NO_3^-	Nitrate		
SO_3^{2-}	Sulfite		
SO_4^{2-}	Sulfate	ClO^-	Hypochlorite
HSO_4^-	Hydrogen sulfate (bisulfate is a widely used common name)	ClO_2^-	Chlorite
		ClO_3^-	Chlorate
		ClO_4^-	Perchlorate
OH^-	Hydroxide	$C_2H_3O_2^-$	Acetate
CN^-	Cyanide	MnO_4^-	Permanganate
PO_4^{3-}	Phosphate	$Cr_2O_7^{2-}$	Dichromate
HPO_4^{2-}	Hydrogen phosphate	CrO_4^{2-}	Chromate
$H_2PO_4^-$	Dihydrogen phosphate	O_2^{2-}	Peroxide
		$C_2O_4^{2-}$	Oxalate

Ionic Compounds with Polyatomic Ions

Polyatomic ion formulas must be memorized.

We have not yet considered ionic compounds that contain polyatomic ions. For example, the compound ammonium nitrate, NH_4NO_3, contains the polyatomic ions NH_4^+ and NO_3^-. Polyatomic ions are assigned special names that *must be memorized* to name the compounds containing them. The most important polyatomic ions and their names are listed in Table 2.5.

Note in Table 2.5 that several series of anions contain an atom of a given element and different numbers of oxygen atoms. These anions are called **oxyanions.** When there are two members in such a series, the name of the one with the smaller number of oxygen atoms ends in *-ite* and the name of the one with the larger number ends in *-ate*—for example, sulfite (SO_3^{2-}) and sulfate (SO_4^{2-}). When more than two oxyanions make up a series, *hypo-* (less than) and *per-* (more than) are used as prefixes to name the members of the series with the fewest and the most oxygen atoms, respectively. The best example involves the oxyanions containing chlorine, as shown in Table 2.5.

Sample Exercise 2.7 Naming Compounds Containing Polyatomic Ions

1. Give the systematic name for each of the following compounds:

a. Na_2SO_4
b. KH_2PO_4
c. $Fe(NO_3)_3$
d. $Mn(OH)_2$
e. Na_2SO_3
f. Na_2CO_3

2. Given the following systematic names, write the formula for each compound:

a. Sodium hydrogen carbonate
b. Cesium perchlorate

c. Sodium hypochlorite
d. Sodium selenate
e. Potassium bromate

Solution

1.

Formula	Name	Comments
a. Na_2SO_4	Sodium sulfate	
b. KH_2PO_4	Potassium dihydrogen phosphate	
c. $Fe(NO_3)_3$	Iron(III) nitrate	Transition metal—name must contain a Roman numeral. The Fe^{3+} ion balances three NO_3^- ions.
d. $Mn(OH)_2$	Manganese(II) hydroxide	Transition metal—name must contain a Roman numeral. The Mn^{2+} ion balances three OH^- ions.
e. Na_2SO_3	Sodium sulfite	
f. Na_2CO_3	Sodium carbonate	

2.

Name	Formula	Comments
a. Sodium hydrogen carbonate	$NaHCO_3$	Often called sodium bicarbonate.
b. Cesium perchlorate	$CsClO_4$	
c. Sodium hypochlorite	NaOCl	
d. Sodium selenate	Na_2SeO_4	Atoms in the same group, like sulfur and selenium, often form similar ions that are named similarly. Thus SeO_4^{2-} is selenate, like SO_4^{2-} (sulfate).
e. Potassium bromate	$KBrO_3$	As above, BrO_3^- is bromate, like ClO_3^- (chlorate).

See Exercises 2.59 and 2.60.

Binary Covalent Compounds (Type III)

In *binary covalent compounds,* the element names follow the same rules as for binary ionic compounds.

Binary covalent compounds are formed between *two nonmetals*. Although these compounds do not contain ions, they are named very similarly to binary ionic compounds.

In the naming of binary covalent compounds, the following rules apply:

1. The first element in the formula is named first, using the full element name.
2. The second element is named as if it were an anion.
3. Prefixes are used to denote the numbers of atoms present. These prefixes are given in Table 2.6.
4. The prefix *mono-* is never used for naming the first element. For example, CO is called carbon monoxide, *not* monocarbon monoxide.

TABLE 2.6 Prefixes Used to Indicate Number in Chemical Names

Prefix	Number Indicated
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

To see how these rules apply, we will now consider the names of the several covalent compounds formed by nitrogen and oxygen:

Compound	Systematic Name	Common Name
N_2O	Dinitrogen monoxide	Nitrous oxide
NO	Nitrogen monoxide	Nitric oxide
NO_2	Nitrogen dioxide	
N_2O_3	Dinitrogen trioxide	
N_2O_4	Dinitrogen tetroxide	
N_2O_5	Dinitrogen pentoxide	

Notice from the preceding examples that to avoid awkward pronunciations, we often drop the final *o* or *a* of the prefix when the element begins with a vowel. For example, N_2O_4 is called dinitrogen tetroxide, *not* dinitrogen tetr*a*oxide, and CO is called carbon monoxide, *not* carbon mon*o*oxide.

Some compounds are always referred to by their common names. The two best examples are water and ammonia. The systematic names for H_2O and NH_3 are never used.

Sample Exercise 2.8

Naming Type III Binary Compounds

1. Name each of the following compounds:

a. PCl_5
b. PCl_3
c. SO_2

2. From the following systematic names, write the formula for each compound:

a. Sulfur hexafluoride
b. Sulfur trioxide
c. Carbon dioxide

Solution

1.

Formula	Name
a. PCl_5	Phosphorus pentachloride
b. PCl_3	Phosphorus trichloride
c. SO_2	Sulfur dioxide

2.

Name	Formula
a. Sulfur hexafluoride	SF_6
b. Sulfur trioxide	SO_3
c. Carbon dioxide	CO_2

See Exercises 2.61 and 2.62.

The rules for naming binary compounds are summarized in Fig. 2.23. Prefixes to indicate the number of atoms are used only in Type III binary compounds (those containing two nonmetals). An overall strategy for naming compounds is given in Fig. 2.24.

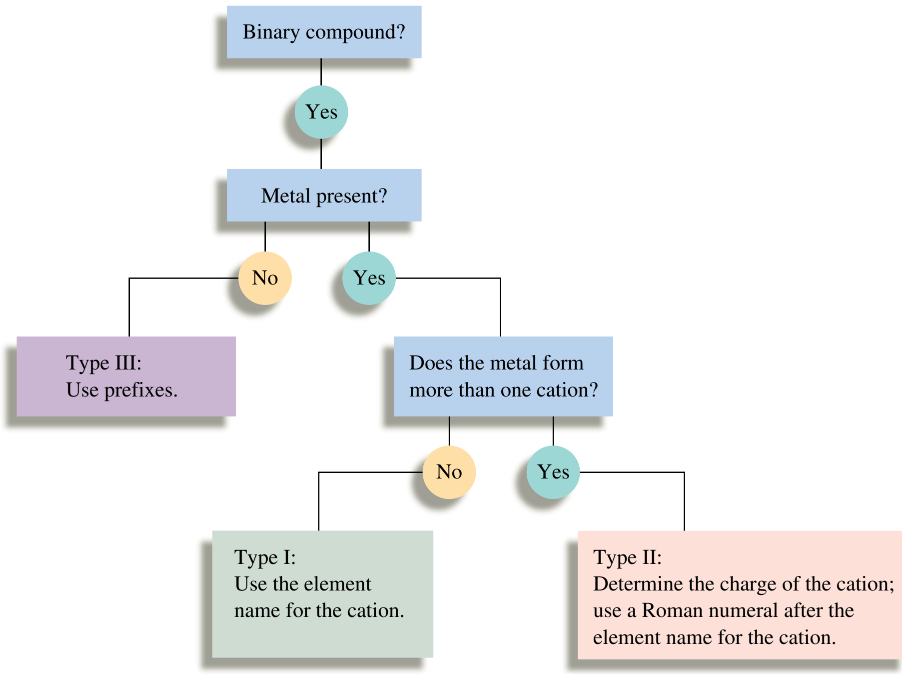


FIGURE 2.23
A flowchart for naming binary compounds.

Sample Exercise 2.9 Naming Various Types of Compounds

1. Give the systematic name for each of the following compounds:

   a. P_4O_{10}
   b. Nb_2O_5
   c. Li_2O_2
   d. $Ti(NO_3)_4$

2. Given the following systematic names, write the formula for each compound:

   a. Vanadium(V) fluoride
   b. Dioxygen difluoride
   c. Rubidium peroxide
   d. Gallium oxide

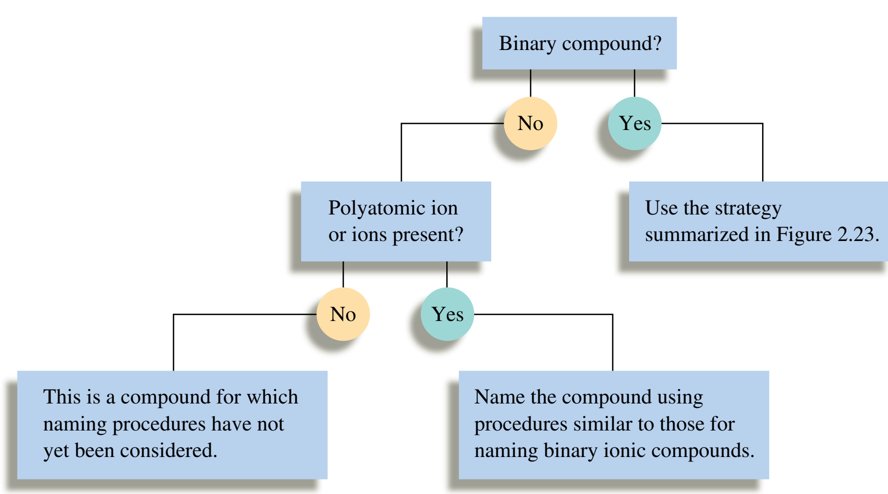


FIGURE 2.24
Overall strategy for naming chemical compounds.

Solution

1.

Compound	Name	Comment
a. P_4O_{10}	Tetraphosphorus decaoxide	Binary covalent compound (Type III), so prefixes are used. The *a* in *deca-* is sometimes dropped.
b. Nb_2O_5	Niobium(V) oxide	Type II binary compound containing Nb^{5+} and O^{2-} ions. Niobium is a transition metal and requires a Roman numeral.
c. Li_2O_2	Lithium peroxide	Type I binary compound containing the Li^+ and O_2^{2-} (peroxide) ions.
d. $Ti(NO_3)_4$	Titanium(IV) nitrate	Not a binary compound. Contains the Ti^{4+} and NO_3^- ions. Titanium is a transition metal and requires a Roman numeral.

2.

Name	Chemical Formula	Comment
a. Vanadium(V) fluoride	VF_5	The compound contains V^{5+} ions and requires five F^- ions for charge balance.
b. Dioxygen difluoride	O_2F_2	The prefix *di-* indicates two of each atom.
c. Rubidium peroxide	Rb_2O_2	Because rubidium is in Group 1A, it forms only 1+ ions. Thus two Rb^+ ions are needed to balance the 2− charge on the peroxide ion (O_2^{2-}).
d. Gallium oxide	Ga_2O_3	Because gallium is in Group 3A, like aluminum, it forms only 3+ ions. Two Ga^{3+} ions are required to balance the charge on three O^{2-} ions.

See Exercises 2.63, 2.65, and 2.66.

Acids

Acids can be recognized by the hydrogen that appears first in the formula.

When dissolved in water, certain molecules produce a solution containing free H^+ ions (protons). These substances, **acids,** will be discussed in detail in Chapters 4, 14, and 15. Here we will simply present the rules for naming acids.

An acid can be viewed as a molecule with one or more H^+ ions attached to an anion. The rules for naming acids depend on whether the anion contains oxygen. If the *anion does not contain oxygen,* the acid is named with the prefix *hydro-* and the suffix *-ic.* For example, when gaseous HCl is dissolved in water, it forms hydrochloric acid. Similarly, HCN and H_2S dissolved in water are called hydrocyanic and hydrosulfuric acids, respectively.

When the *anion contains oxygen,* the acidic name is formed from the root name of the anion with a suffix of *-ic* or *-ous,* depending on the name of the anion.

1. If the anion name ends in *-ate,* the suffix *-ic* is added to the root name. For example, H_2SO_4 contains the sulfate anion (SO_4^{2-}) and is called sulfuric acid; H_3PO_4 contains the phosphate anion (PO_4^{3-}) and is called phosphoric acid; and $HC_2H_3O_2$ contains the acetate ion ($C_2H_3O_2^-$) and is called acetic acid.
2. If the anion has an *-ite* ending, the *-ite* is replaced by *-ous.* For example, H_2SO_3, which contains sulfite (SO_3^{2-}), is named sulfurous acid; and HNO_2, which contains nitrite (NO_2^-), is named nitrous acid.

TABLE 2.7 Names of Acids* That Do Not Contain Oxygen

Acid	Name
HF	Hydrofluoric acid
HCl	Hydrochloric acid
HBr	Hydrobromic acid
HI	Hydroiodic acid
HCN	Hydrocyanic acid
H_2S	Hydrosulfuric acid

*Note that these acids are aqueous solutions containing these substances.

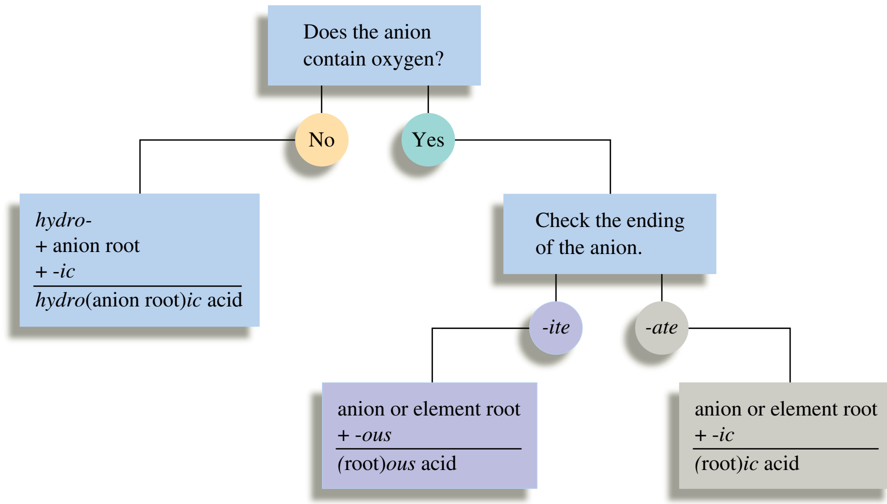


FIGURE 2.25
A flowchart for naming acids. An acid is best considered as one or more H^+ ions attached to an anion.

TABLE 2.8 Names of Some Oxygen-Containing Acids

Acid	Name
HNO_3	Nitric acid
HNO_2	Nitrous acid
H_2SO_4	Sulfuric acid
H_2SO_3	Sulfurous acid
H_3PO_4	Phosphoric acid
$HC_2H_3O_2$	Acetic Acid

The application of these rules can be seen in the names of the acids of the oxyanions of chlorine:

Acid	Anion	Name
$HClO_4$	Perchlor*ate*	Perchlor*ic* acid
$HClO_3$	Chlor*ate*	Chlor*ic* acid
$HClO_2$	Chlor*ite*	Chlor*ous* acid
$HClO$	Hypochlor*ite*	Hypochlor*ous* acid

The names of the most important acids are given in Tables 2.7 and 2.8. An overall strategy for naming acids is shown in Fig. 2.25.

Key Terms

Section 2.2
law of conservation of mass
law of definite proportion
law of multiple proportions

Section 2.3
atomic masses
atomic weights
Avogadro's hypothesis

Section 2.4
cathode-ray tube
electron
radioactivity
nuclear atom
nucleus

Section 2.5
proton
neutron
isotopes
atomic number
mass number

For Review

Fundamental laws
- Conservation of mass
- Definite proportion
- Multiple proportions

Dalton's atomic theory
- All elements are composed of atoms.
- All atoms of a given element are identical.
- Chemical compounds are formed when atoms combine.
- Atoms are not changed in chemical reactions but the way they are bound together changes.

Early atomic experiments and models
- Thomson model
- Millikan experiment
- Rutherford experiment
- Nuclear model

Section 2.6
chemical bond
covalent bond
molecule
chemical formula
structural formula
space-filling model
ball-and-stick model
ion
cation
anion
ionic bond
ionic solid (salt)
polyatomic ion

Section 2.7
periodic table
metal
nonmetal
group (family)
alkali metals
alkaline earth metals
halogens
noble gases
period

Section 2.8
binary compounds
binary ionic compounds
oxyanions
binary covalent compounds
acid

Atomic structure
- Small dense nucleus contains protons and neutrons.
  - Protons—positive charge
  - Neutrons—no charge
- Electrons reside outside the nucleus in the relatively large remaining atomic volume.
  - Electrons—negative charge, small mass (1/1840 of proton)
- Isotopes have the same atomic number but different mass numbers.

Atoms combine to form molecules by sharing electrons to form covalent bonds.
- Molecules are described by chemical formulas.
- Chemical formulas show number and type of atoms.
  - Structural formula
  - Ball-and-stick model
  - Space-filling model

Formation of ions
- Cation—formed by loss of an electron, positive charge
- Anion—formed by gain of an electron, negative charge
- Ionic bonds—formed by interaction of cations and anions

The periodic table organizes elements in order of increasing atomic number.
- Elements with similar properties are in columns, or groups.
- Metals are in the majority and tend to form cations.
- Nonmetals tend to form anions.

Compounds are named using a system of rules depending on the type of compound.
- Binary compounds
  - Type I—contain a metal that always forms the same cation
  - Type II—contain a metal that can form more than one cation
  - Type III—contain two nonmetals
- Compounds containing a polyatomic ion

REVIEW QUESTIONS

1. Use Dalton's atomic theory to account for each of the following.
   a. the law of conservation of mass
   b. the law of definite proportion
   c. the law of multiple proportions
2. What evidence led to the conclusion that cathode rays had a negative charge?
3. What discoveries were made by J. J. Thomson, Henri Becquerel, and Lord Rutherford? How did Dalton's model of the atom have to be modified to account for these discoveries?
4. Consider Ernest Rutherford's alpha-particle bombardment experiment illustrated in Figure 2.12. How did the results of this experiment lead Rutherford away from the plum pudding model of the atom to propose the nuclear model of the atom?
5. Do the proton and the neutron have exactly the same mass? How do the masses of the proton and neutron compare to the mass of the electron? Which particles make the greatest contribution to the mass of an atom? Which particles make the greatest contribution to the chemical properties of an atom?
6. What is the distinction between atomic number and mass number? Between mass number and atomic mass?
7. Distinguish between the terms *family* and *period* in connection with the periodic table. For which of these terms is the term *group* also used?
8. The compounds $AlCl_3$, $CrCl_3$, and ICl_3 have similar formulas, yet each follows a different set of rules to name it. Name these compounds, and then compare and contrast the nomenclature rules used in each case.

9. When metals react with nonmetals, an ionic compound generally results. What is the predicted general formula for the compound formed between an alkali metal and sulfur? Between an alkaline earth metal and nitrogen? Between aluminum and a halogen?
10. How would you name $HBrO_4$, KIO_3, $NaBrO_2$, and HIO? Refer to Table 2.5 and the acid nomenclature discussion in the text.

Active Learning Questions

These questions are designed to be used by groups of students in class. The questions allow students to explore their understanding of concepts through discussion and peer teaching. The real value of these questions is the learning that occurs while students talk to each other about chemical concepts.

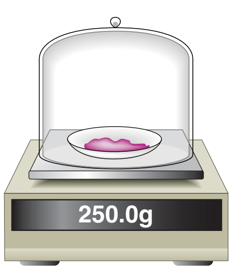
1. Which of the following is true about an individual atom? Explain.
   a. An individual atom should be considered to be a solid.
   b. An individual atom should be considered to be a liquid.
   c. An individual atom should be considered to be a gas.
   d. The state of the atom depends on which element it is.
   e. An individual atom cannot be considered to be a solid, liquid, or gas.
   Justify your choice, and for choices you did not pick, explain what is wrong with them.

2. How would you go about finding the number of "chalk molecules" it takes to write your name on the board? Provide an explanation of all you would need to do and a sample calculation.

3. These questions concern the work of J. J. Thomson.
   a. From Thomson's work, which particles do you think he would feel are most important for the formation of compounds (chemical changes) and why?
   b. Of the remaining two subatomic particles, which do you place second in importance for forming compounds and why?
   c. Propose three models that explain Thomson's findings and evaluate them. To be complete you should include Thomson's findings.

4. Heat is applied to an ice cube in a closed container until only steam is present. Draw a representation of this process, assuming you can see it at an extremely high level of magnification. What happens to the size of the molecules? What happens to the total mass of the sample?

5. You have a chemical in a sealed glass container filled with air. The setup is sitting on a balance as shown below. The chemical is ignited by means of a magnifying glass focusing sunlight on the reactant. After the chemical has completely burned, which of the following is true? Explain your answer.



   a. The balance will read less than 250.0 g.
   b. The balance will read 250.0 g.
   c. The balance will read greater than 250.0 g.
   d. Cannot be determined without knowing the identity of the chemical.

6. You take three compounds consisting of two elements and decompose them. To determine the relative masses of *X*, *Y*, and *Z*, you collect and weigh the elements, obtaining the following data:

Elements in Compound	Masses of Elements
X and *Y*	$X = 0.4$ g, $Y = 4.2$ g
Y and *Z*	$Y = 1.4$ g, $Z = 1.0$ g
X and *Y*	$X = 2.0$ g, $Y = 7.0$ g

   a. What are the assumptions in solving this problem?
   b. What are the relative masses of *X*, *Y*, and *Z*?
   c. What are the chemical formulas of the three compounds?
   d. If you decompose 21 g of compound *XY*, how much of each element is present?

7. The vitamin niacin (nicotinic acid, $C_6H_5NO_2$) can be isolated from a variety of natural sources such as liver, yeast, milk, and whole grain. It also can be synthesized from commercially available materials. Which source of nicotinic acid, from a nutritional view, is best for use in a multivitamin tablet? Why?

8. One of the best indications of a useful theory is that it raises more questions for further experimentation than it originally answered. Does this apply to Dalton's atomic theory? Give examples.

9. Dalton assumed that all atoms of the same element were identical in all their properties. Explain why this assumption is not valid.

10. Evaluate each of the following as an acceptable name for water:
   a. dihydrogen oxide
   b. hydroxide hydride
   c. hydrogen hydroxide
   d. oxygen dihydride

11. Why do we call $Ba(NO_3)_2$ barium nitrate, but we call $Fe(NO_3)_2$ iron(II) nitrate?

12. Why is calcium dichloride not the correct systematic name for $CaCl_2$?

13. The common name for NH_3 is ammonia. What would be the systematic name for NH_3? Support your answer.

A blue question or exercise number indicates that the answer to that question or exercise appears at the back of this book and a solution appears in the *Solutions Guide*.

Questions

14. What refinements had to be made in Dalton's atomic theory to account for Gay-Lussac's results on the combining volumes of gases?

15. When hydrogen is burned in oxygen to form water, the composition of water formed does not depend on the amount of oxygen reacted. Interpret this in terms of the law of definite proportion.
16. The two most reactive families of elements are the halogens and the alkali metals. How do they differ in their reactivities?
17. Explain the law of conservation of mass, the law of definite proportion, and the law of multiple proportions.
18. Section 2.3 describes the postulates of Dalton's atomic theory. With some modifications, these postulates hold up very well regarding how we view elements, compounds, and chemical reactions today. Answer the following questions concerning Dalton's atomic theory and the modifications made today.
    a. The atom can be broken down into smaller parts. What are the smaller parts?
    b. How are atoms of hydrogen identical to each other and how can they be different from each other?
    c. How are atoms of hydrogen different from atoms of helium? How can H atoms be similar to He atoms?
    d. How is water different from hydrogen peroxide (H_2O_2) even though both compounds are composed of only hydrogen and oxygen?
    e. What happens in a chemical reaction and why is mass conserved in a chemical reaction?
19. The contributions of J. J. Thomson and Ernest Rutherford led the way to today's understanding of the structure of the atom. What were their contributions?
20. What is the modern view of the structure of the atom?
21. The number of protons in an atom determines the identity of the atom. What does the number and arrangement of the electrons in an atom determine? What does the number of neutrons in an atom determine?
22. Distinguish between the following terms.
    a. molecule versus ion
    b. covalent bonding versus ionic bonding
    c. molecule versus compound
    d. anion versus cation
23. Which of the following statements are true? For the false statements, correct them.
    a. Most of the known elements are metals.
    b. Element 118 should be a nonmetal.
    c. Hydrogen has mostly metallic properties.
    d. A family of elements is also known as a period of elements.
    e. When an alkaline earth metal, A, reacts with a halogen, X, the formula of the covalent compound formed should be A_2X.
24. Each of the following compounds has three possible names listed for it. For each compound, what is the correct name and why aren't the other names used?
    a. N_2O: nitrogen oxide, nitrogen(I) oxide, dinitrogen monoxide
    b. Cu_2O: copper oxide, copper(I) oxide, dicopper monoxide
    c. Li_2O: lithium oxide, lithium(I) oxide, dilithium monoxide

Exercises

In this section similar exercises are paired.

Development of the Atomic Theory

25. When mixtures of gaseous H_2 and gaseous Cl_2 react, a product forms that has the same properties regardless of the relative amounts of H_2 and Cl_2 used.
    a. How is this result interpreted in terms of the law of definite proportion?
    b. When a volume of H_2 reacts with an equal volume of Cl_2 at the same temperature and pressure, what volume of product having the formula HCl is formed?
26. A reaction of 1 liter of chlorine gas (Cl_2) with 3 liters of fluorine gas (F_2) yields 2 liters of a gaseous product. All gas volumes are at the same temperature and pressure. What is the formula of the gaseous product?

27. Hydrazine, ammonia, and hydrogen azide all contain only nitrogen and hydrogen. The mass of hydrogen that combines with 1.00 g of nitrogen for each compound is 1.44×10^{-1} g, 2.16×10^{-1} g, and 2.40×10^{-2} g, respectively. Show how these data illustrate the law of multiple proportions.
28. Consider 100.0-g samples of two different compounds consisting only of carbon and oxygen. One compound contains 27.2 g of carbon and the other has 42.9 g of carbon. How can these data support the law of multiple proportions if 42.9 is not a multiple of 27.2? Show that these data support the law of multiple proportions.

29. Early tables of atomic weights (masses) were generated by measuring the mass of a substance that reacts with 1.00 g of oxygen. Given the following data and taking the atomic mass of hydrogen as 1.00, generate a table of relative atomic masses for oxygen, sodium, and magnesium.

Element	Mass That Combines with 1.00 g Oxygen	Assumed Formula
Hydrogen	0.126 g	HO
Sodium	2.875 g	NaO
Magnesium	1.500 g	MgO

How do your values compare with those in the periodic table? How do you account for any differences?

30. Indium oxide contains 4.784 g of indium for every 1.000 g of oxygen. In 1869, when Mendeleev first presented his version of the periodic table, he proposed the formula In_2O_3 for indium oxide. Before that time it was thought that the formula was InO. What values for the atomic mass of indium are obtained using these two formulas? Assume that oxygen has an atomic mass of 16.00.

The Nature of the Atom

31. From the information in this chapter on the mass of the proton, the mass of the electron, and the sizes of the nucleus and the atom, calculate the densities of a hydrogen nucleus and a hydrogen atom.
32. If you wanted to make an accurate scale model of the hydrogen atom and decided that the nucleus would have a diameter of 1 mm, what would be the diameter of the entire model?

33. In an experiment it was found that the total charge on an oil drop was 5.93×10^{-18} C. How many negative charges does the drop contain?
34. A chemist in a galaxy far, far away performed the Millikan oil drop experiment and got the following results for the charges on

various drops. Use these data to calculate the charge of the electron in zirkombs.

2.56×10^{-12} zirkombs	7.68×10^{-12} zirkombs
3.84×10^{-12} zirkombs	6.40×10^{-13} zirkombs

35. What are the symbols of the following metals: sodium, radium, iron, gold, manganese, lead.

36. What are the symbols of the following nonmetals: fluorine, chlorine, bromine, sulfur, oxygen, phosphorus?

37. Give the names of the metals that correspond to the following symbols: Sn, Pt, Hg, Mg, K, Ag.

38. Give the names of the nonmetals that correspond to the following symbols: As, I, Xe, He, C, Si.

39. **a.** Classify the following elements as metals or nonmetals:

Mg	Si	Rn
Ti	Ge	Eu
Au	B	Am
Bi	At	Br

b. The distinction between metals and nonmetals is really not a clear one. Some elements, called *metalloids,* are intermediate in their properties. Which of these elements would you reclassify as metalloids? What other elements in the periodic table would you expect to be metalloids?

40. **a.** List the noble gas elements. Which of the noble gases has only radioactive isotopes? (This situation is indicated on most periodic tables by parentheses around the mass of the element. See inside front cover.)

b. Which lanthanide element and which transition element have only radioactive isotopes?

41. In the periodic table, how many elements are found in

a. Group 2A?
b. the oxygen family?
c. the nickel group?
d. Group 8A?

42. In the periodic table, how many elements are found

a. in the halogen group?
b. in the alkali family?
c. in the lanthanide series?
d. classified as transition metals?

43. How many protons and neutrons are in the nucleus of each of the following atoms? In a neutral atom of each element, how many electrons are present?

a. ^{79}Br
b. ^{81}Br
c. ^{239}Pu
d. ^{133}Cs
e. ^{3}H
f. ^{56}Fe

44. What number of protons and neutrons are contained in the nucleus of each of the following atoms? Assuming each atom is uncharged, what number of electrons are present?

a. $^{235}_{92}U$
b. $^{13}_{6}C$
c. $^{57}_{26}Fe$
d. $^{208}_{82}Pb$
e. $^{86}_{37}Rb$
f. $^{41}_{20}Ca$

45. Write the atomic symbol ($^{A}_{Z}X$) for each of the following isotopes.

a. $Z = 8$, number of neutrons = 9
b. the isotope of chlorine in which $A = 37$
c. $Z = 27$, $A = 60$
d. number of protons = 26, number of neutrons = 31
e. the isotope of I with a mass number of 131
f. $Z = 3$, number of neutrons = 4

46. Write the atomic symbol ($^{A}_{Z}X$) for each of the isotopes described below.

a. number of protons = 27, number of neutrons = 31
b. the isotope of boron with mass number 10
c. $Z = 12$, $A = 23$
d. atomic number 53, number of neutrons = 79
e. $Z = 9$, number of neutrons = 10
f. number of protons = 29, mass number 65

47. What is the symbol for an ion with 63 protons, 60 electrons, and 88 neutrons? If an ion contains 50 protons, 68 neutrons, and 48 electrons, what is its symbol?

48. What is the symbol of an ion with 16 protons, 18 neutrons, and 18 electrons? What is the symbol for an ion that has 16 protons, 16 neutrons, and 18 electrons?

49. Complete the following table:

Symbol	Number of Protons in Nucleus	Number of Neutrons in Nucleus	Number of Electrons	Net Charge
$^{238}_{92}U$				
	20	20		2+
	23	28	20	
$^{89}_{39}Y$				
	35	44	36	
	15	16		3−

50.

Symbol	Number of Protons in Nucleus	Number of Neutrons in Nucleus	Number of Electrons	Net Charge
$^{53}_{26}Fe^{2+}$				
	26	33		3+
	85	125	86	
	13	14	10	
		76	54	2−

51. For each of the following sets of elements, label each as either noble gases, halogens, alkali metals, alkaline earth metals, or transition metals.

a. Ti, Fe, Ag
b. Mg, Sr, Ba
c. Li, K, Rb
d. Ne, Kr, Xe
e. F, Br, I

52. Consider the elements of Group 4A (the "carbon family"): C, Si, Ge, Sn, and Pb. What is the trend in metallic character as one goes down this group? What is the trend in metallic character going from left to right across a period in the periodic table?

53. Would you expect each of the following atoms to gain or lose electrons when forming ions? What ion is the most likely in each case?

a. Ra **c.** P **e.** Br
b. In **d.** Te **f.** Rb

54. For each of the following atomic numbers, use the periodic table to write the formula (including the charge) for the simple *ion* that the element is most likely to form in ionic compounds.

a. 13 **c.** 56 **e.** 87
b. 34 **d.** 7 **f.** 35

Nomenclature

55. Name the compounds in parts a–d and write the formulas for the compounds in parts e–h.

a. NaBr **e.** strontium fluoride
b. Rb_2O **f.** aluminum selenide
c. CaS **g.** potassium nitride
d. AlI_3 **h.** magnesium phosphide

56. Name the compounds in parts a–d and write the formulas for the compounds in parts e–h.

a. Hg_2O **e.** tin(II) nitride
b. $FeBr_3$ **f.** cobalt(III) iodide
c. CoS **g.** mercury(II) oxide
d. $TiCl_4$ **h.** chromium(VI) sulfide

57. Name each of the following compounds:

a. CsF **c.** Ag_2S **e.** TiO_2
b. Li_3N **d.** MnO_2 **f.** Sr_3P_2

58. Write the formula for each of the following compounds:

a. zinc chloride **d.** aluminum sulfide
b. tin(IV) fluoride **e.** mercury(I) selenide
c. calcium nitride **f.** silver iodide

59. Name each of the following compounds:

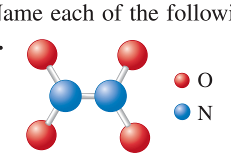
a. $BaSO_3$ **c.** $KMnO_4$
b. $NaNO_2$ **d.** $K_2Cr_2O_7$

60. Write the formula for each of the following compounds:

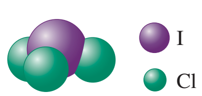
a. chromium(III) hydroxide **c.** lead(IV) carbonate
b. magnesium cyanide **d.** ammonium acetate

61. Name each of the following compounds:

a.



b.



c. SO_2
d. P_2S_5

62. Write the formula for each of the following compounds:

a. diboron trioxide **c.** dinitrogen monoxide
b. arsenic pentafluoride **d.** sulfur hexachloride

63. Name each of the following compounds:

a. CuI **c.** CoI_2
b. CuI_2 **d.** Na_2CO_3
e. $NaHCO_3$ **h.** NaOCl
f. S_4N_4 **i.** $BaCrO_4$
g. SF_6 **j.** NH_4NO_3

64. Name each of the following compounds:

a. $HC_2H_3O_2$ **g.** H_2SO_4
b. NH_4NO_2 **h.** Sr_3N_2
c. Co_2S_3 **i.** $Al_2(SO_3)_3$
d. ICl **j.** SnO_2
e. $Pb_3(PO_4)_2$ **k.** Na_2CrO_4
f. KIO_3 **l.** HClO

65. Write the formula for each of the following compounds:

a. sulfur difluoride
b. sulfur hexafluoride
c. sodium dihydrogen phosphate
d. lithium nitride
e. chromium(III) carbonate
f. tin(II) fluoride
g. ammonium acetate
h. ammonium hydrogen sulfate
i. cobalt(III) nitrate
j. mercury(I) chloride
k. potassium chlorate
l. sodium hydride

66. Write the formula for each of the following compounds:

a. chromium(VI) oxide
b. disulfur dichloride
c. nickel(II) fluoride
d. potassium hydrogen phosphate
e. aluminum nitride
f. ammonia
g. manganese(IV) sulfide
h. sodium dichromate
i. ammonium sulfite
j. carbon tetraiodide

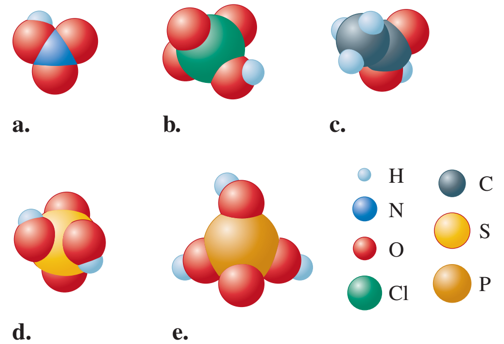
67. Write the formula for each of the following compounds:

a. sodium oxide **h.** copper(I) chloride
b. sodium peroxide **i.** gallium arsenide
c. potassium cyanide **j.** cadmium selenide
d. copper(II) nitrate **k.** zinc sulfide
e. selenium tetrabromide **l.** nitrous acid
f. iodous acid **m.** diphosphorus pentoxide
g. lead(IV) sulfide

68. Write the formula for each of the following compounds:

a. ammonium hydrogen phosphate
b. mercury(I) sulfide
c. silicon dioxide
d. sodium sulfite
e. aluminum hydrogen sulfate
f. nitrogen trichloride
g. hydrobromic acid
h. bromous acid
i. perbromic acid
j. potassium hydrogen sulfide
k. calcium iodide
l. cesium perchlorate

69. Name the following acids illustrated below.



70. Each of the following compounds is incorrectly named. What is wrong with each name, and what is the correct name for each compound?
   a. $FeCl_3$, iron chloride
   b. NO_2, nitrogen(IV) oxide
   c. CaO, calcium(II) monoxide
   d. Al_2S_3, dialuminum trisulfide
   e. $Mg(C_2H_3O_2)_2$, manganese diacetate
   f. $FePO_4$, iron(II) phosphide
   g. P_2S_5, phosphorous sulfide
   h. Na_2O_2, sodium oxide
   i. HNO_3, nitrate acid
   j. H_2S, sulfuric acid

Additional Exercises

71. Chlorine has two natural isotopes: $^{37}_{17}Cl$ and $^{35}_{17}Cl$. Hydrogen reacts with chlorine to form the compound HCl. Would a given amount of hydrogen react with different masses of the two chlorine isotopes? Does this conflict with the law of definite proportion? Why or why not?

72. Which of the following statements is(are) *true?* For the false statements, correct them.
   a. All particles in the nucleus of an atom are charged.
   b. The atom is best described as a uniform sphere of matter in which electrons are embedded.
   c. The mass of the nucleus is only a very small fraction of the mass of the entire atom.
   d. The volume of the nucleus is only a very small fraction of the total volume of the atom.
   e. The number of neutrons in a neutral atom must equal the number of electrons.

73. The isotope of an unknown element, X, has a mass number of 79. The most stable ion of the isotope has 36 electrons and forms a binary compound with sodium having a formula of Na_2X. Which of the following statements is(are) *true?* For the false statements, correct them.
   a. The binary compound formed between X and fluorine will be a covalent compound.
   b. The isotope of X contains 38 protons.
   c. The isotope of X contains 41 neutrons.
   d. The identity of X is strontium, Sr.

74. For each of the following ions, indicate the total number of protons and electrons in the ion. For the positive ions in the list, predict the formula of the simplest compound formed between each positive ion and the oxide ion. For the negative ions in the list, predict the formula of the simplest compound formed between each negative ion and the aluminum ion.

a. Fe^{2+}	e. S^{2-}
b. Fe^{3+}	f. P^{3-}
c. Ba^{2+}	g. Br^{-}
d. Cs^{+}	h. N^{3-}

75. The formulas and common names for several substances are given below. Give the systematic names for these substances.

a. sugar of lead	$Pb(C_2H_3O_2)_2$
b. blue vitrol	$CuSO_4$
c. quicklime	CaO
d. Epsom salts	$MgSO_4$
e. milk of magnesia	$Mg(OH)_2$
f. gypsum	$CaSO_4$
g. laughing gas	N_2O

76. Identify each of the following elements:
   a. a member of the same family as oxygen whose most stable ion contains 54 electrons
   b. a member of the alkali metal family whose most stable ion contains 36 electrons
   c. a noble gas with 18 protons in the nucleus
   d. a halogen with 85 protons and 85 electrons

77. An element's most stable ion forms an ionic compound with bromine, having the formula XBr_2. If the ion of element X has a mass number of 230 and has 86 electrons, what is the identity of the element, and how many neutrons does it have?

78. A certain element has only two naturally occurring isotopes: one with 18 neutrons and the other with 20 neutrons. The element forms 1− charged ions when in ionic compounds. Predict the identity of the element. What number of electrons does the 1− charged ion have?

79. The designations 1A through 8A used for certain families of the periodic table are helpful for predicting the charges on ions in binary ionic compounds. In these compounds, the metals generally take on a positive charge equal to the family number, while the nonmetals take on a negative charge equal to the family number minus eight. Thus the compound between sodium and chlorine contains Na^+ ions and Cl^- ions and has the formula NaCl. Predict the formula and the name of the binary compound formed from the following pairs of elements.

a. Ca and N	e. Ba and I
b. K and O	f. Al and Se
c. Rb and F	g. Cs and P
d. Mg and S	h. In and Br

80. By analogy with phosphorous compounds, name the following: Na_3AsO_4, H_3AsO_4, $Mg_3(SbO_4)_2$.

81. A sample of H_2SO_4 contains 2.02 g of hydrogen, 32.07 g of sulfur, and 64.00 g of oxygen. How many grams of sulfur and grams of oxygen are present in a second sample of H_2SO_4 containing 7.27 g of hydrogen?

82. In a reaction, 34.0 g of chromium(III) oxide reacts with 12.1 g of aluminum to produce chromium and aluminum oxide. If 23.3 g of chromium is produced, what mass of aluminum oxide is produced?

Challenge Problems

83. The elements in one of the groups in the periodic table are often called the coinage metals. Identify the elements in this group based on your own experience.

84. Reaction of 2.0 L of hydrogen gas with 1.0 L of oxygen gas yields 2.0 L of water vapor. All gases are at the same temperature and pressure. Show how these data support the idea that oxygen gas is a diatomic molecule. Must we consider hydrogen to be a diatomic molecule to explain these results?

85. A combustion reaction involves the reaction of a substance with oxygen gas. The complete combustion of any hydrocarbon (binary compound of carbon and hydrogen) produces carbon dioxide and water as the only products. Octane is a hydrocarbon that is found in gasoline. Complete combustion of octane produces 8 liters of carbon dioxide for every 9 liters of water vapor (both measured at the same temperature and pressure). What is the ratio of carbon atoms to hydrogen atoms in a molecule of octane?

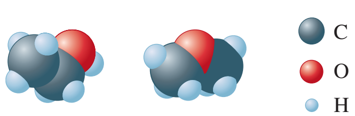
86. A chemistry instructor makes the following claim: "Consider that if the nucleus were the size of a grape, the electrons would be about 1 *mile* away on average." Is this claim reasonably accurate? Provide mathematical support.

87. Two elements, R and Q, combine to form two binary compounds. In the first compound, 14.0 g of R combines with 3.00 g of Q. In the second compound, 7.00 g of R combines with 4.50 g of Q. Show that these data are in accord with the law of multiple proportions. If the formula of the second compound is RQ, what is the formula of the first compound?

88. The early alchemists used to do an experiment in which water was boiled for several days in a sealed glass container. Eventually, some solid residue would appear in the bottom of the flask, which was interpreted to mean that some of the water in the flask had been converted into "earth." When Lavoisier repeated this experiment, he found that the water weighed the same before and after heating and the mass of the flask plus the solid residue equaled the original mass of the flask. Were the alchemists correct? Explain what really happened. (This experiment is described in the article by A. F. Scott in *Scientific American,* January 1984.)

89. Each of the following statements is true, but Dalton might have had trouble explaining some of them with his atomic theory. Give explanations for the following statements.

   a. The space-filling models for ethyl alcohol and dimethyl ether are shown below.



   These two compounds have the same composition by mass (52% carbon, 13% hydrogen, and 35% oxygen), yet the two have different melting points, boiling points, and solubilities in water.

   b. Burning wood leaves an ash that is only a small fraction of the mass of the original wood.

   c. Atoms can be broken down into smaller particles.

   d. One sample of lithium hydride is 87.4% lithium by mass, while another sample of lithium hydride is 74.9% lithium by mass. However, the two samples have the same properties.

90. You have two distinct gaseous compounds made from element X and element Y. The mass percents are as follows:

   Compound I: 30.43% X, 69.57% Y
   Compound II: 63.64% X, 36.36% Y

   In their natural standard states, element X and element Y exist as gases. (Monatomic? Diatomic? Triatomic? That is for you to determine.) When you react "gas X" with "gas Y" to make the products, you get the following data (all at standard pressure and temperature):

   1 volume "gas X" + 2 volumes "gas Y" $\longrightarrow$ 2 volumes compound I

   2 volumes "gas X" + 1 volume "gas Y" $\longrightarrow$ 2 volumes compound II

   Assume the simplest possible formulas for reactants and products in the chemical equations above. Then, determine the relative atomic masses of element X and element Y.

Integrative Problems

These problems require the integration of multiple concepts to find the solutions.

91. What is the systematic name of Ta_2O_5? If the charge on the metal remained constant and then sulfur was substituted for oxygen, how would the formula change? What is the difference in the total number of protons between Ta_2O_5 and its sulfur analog?

92. A binary ionic compound is known to contain a cation with 51 protons and 48 electrons. The anion contains one-third the number of protons as the cation. The number of electrons in the anion is equal to the number of protons plus 1. What is the formula of this compound? What is the name of this compound?

93. Using the information in Table 2.1, answer the following questions. In an ion with an unknown charge, the total mass of all the electrons was determined to be 2.55×10^{-26} g, while the total mass of its protons was 5.34×10^{-23} g. What is the identity and charge of this ion? What is the symbol and mass number of a neutral atom whose total mass of its electrons is 3.92×10^{-26} g, while its neutrons have a mass of 9.35×10^{-23} g?

Marathon Problem

This problem is designed to incorporate several concepts and techniques into one situation. Marathon Problems can be used in class by groups of students to help facilitate problem-solving skills.

94. You have gone back in time and are working with Dalton on a table of relative masses. Following are his data.

   0.602 g gas A reacts with 0.295 g gas B
   0.172 g gas B reacts with 0.401 g gas C
   0.320 g gas A reacts with 0.374 g gas C

   a. Assuming simplest formulas (AB, BC, and AC), construct a table of relative masses for Dalton.

b. Knowing some history of chemistry, you tell Dalton that if he determines the volumes of the gases reacted at constant temperature and pressure, he need not assume simplest formulas. You collect the following data:

6 volumes gas A + 1 volume gas B → 4 volumes product
1 volume gas B + 4 volumes gas C → 4 volumes product
3 volumes gas A + 2 volumes gas C → 6 volumes product

Write the simplest balanced equations, and find the actual relative masses of the elements. Explain your reasoning.

Get help understanding core concepts and visualizing molecular-level interactions, and practice problem solving, by visiting the Online Study Center at **college.hmco.com/PIC/zumdahl7e.**