1 FUNDAMENTAL CONCEPTS AND UNITS OF MEASUREMENT

Mike Peterson #54 of the Jacksonville Jaguars sacks Houston quarterback David Carr #8. Athletes such as these find their sports safer than ever before thanks to high tech materials made possible through chemical research. Most of the materials used in their uniforms, helmets, and protective pads do not occur naturally in our world and would not exist without discoveries made by observant chemists. The fruits of chemical science touch all our lives every day in ways most of us rarely think of. *(Lisa Blumenfeld/Getty Images/NewsCom.)*

CHAPTER OUTLINE

THIS CHAPTER IN CONTEXT Chemistry is a science that has impacted every aspect of our lives. We have come to take for granted so many of the materials, discovered by chemists, that make us comfortable, provide for our entertainment, and ensure that the foods we place on our tables are fresh and wholesome. Most of the medicines to cure disease and relieve pain, and nearly all of the objects used by doctors in hospitals, would not exist if chemists had not synthesized the materials from which they are made. As we guide you through the study of chemistry, we will provide numerous examples of how this subject relates to the world in which we live. Our aim is to give you an appreciation of the significant role that chemistry plays in modern society.

This chapter has three principal goals. The first is to provide you with an appreciation of the central role that chemistry plays among the sciences. The second is to have you understand the way scientists approach the study of nature and how they construct mental pictures of the microscopic world to explain the results of experimental observations. And third, we will begin to discuss the principal substances that serve as building blocks for all the materials we encounter in our daily lives.

If you've had a prior course in chemistry, perhaps in high school, you're likely to be familiar with many of the topics that we cover in this chapter. Nevertheless, it is important to be sure you have a mastery of these subjects, because if you don't start this course with a firm understanding of the basics, you may find yourself in trouble later on.

1.1 CHEMISTRY IS IMPORTANT FOR ANYONE STUDYING THE SCIENCES

In our discussions, we do not assume that you have had a prior course in chemistry. However, we do urge you to study this chapter thoroughly, because the concepts developed here will be used in later chapters.

Chemistry[1] is the study of the composition and properties of matter, which includes all of the chemicals that make up tangible things, from rocks to people to pizza. Chemists search for answers to fundamental questions about the effect of a substance's composition on its properties. They also seek to learn the way substances change, often dramatically, when they interact with each other in *chemical reactions*. And permeating all of this is a search for knowledge about the basic underlying structure of matter and the forces that determine the properties we observe through our senses. From these studies has come the ability to create materials never before found on earth, materials with especially desirable properties that fulfill specific needs of society. This knowledge has also enabled biologists to develop a fundamental understanding of many of the processes taking place in living organisms.

Although you may not plan to be a chemist, some knowledge of chemistry will surely be valuable to you. In fact, the involvement of chemistry among the various branches of science is evidenced by the names of some of the divisions of the American Chemical Society, the largest scientific organization in the world (see Table 1.1).

TABLE 1.1 Names of Some of the Divisions of the American Chemical Society

Agricultural & Food Chemistry	Computers in Chemistry
Agrochemicals	Environmental Chemistry
Biochemical Technology	Fuel Chemistry
Biological Chemistry	Geochemistry
Business Development & Management	Industrial & Engineering Chemistry
Carbohydrate Chemistry	Medicinal Chemistry
Cellulose and Renewable Materials	Nuclear Chemistry & Technology
Chemical Health & Safety	Petroleum Chemistry
Chemical Toxicology	Polymer Chemistry
Chemistry & the Law	Polymeric Materials: Science & Engineering
Colloid & Surface Chemistry	Rubber

[1] Important terms will be set in bold type to call them to your attention. Be sure you learn their meanings.

1.2 The scientific method helps us build models of nature

Scientists who work in university, industrial, and government laboratories follow a general approach to their work called the **scientific method.** In very simple terms, it is a cyclical process in which we gather and assemble information about nature, formulate explanations for what we've observed, and then test the explanations with new experiments.

In the sciences, we usually gather information by performing experiments in laboratories under controlled conditions so observations we make are reproducible (Figure 1.1). An **observation** *is a statement that accurately describes something we see, hear, taste, feel, or smell.*

Observations gathered during an experiment often lead us to make conclusions. A **conclusion** *is a statement that's based on what we think about a series of observations.* For example, consider the following statements about the fermentation of grape juice to make wine:

1. Before fermentation, grape juice is very sweet and contains no alcohol.
2. After fermentation, the grape juice is no longer as sweet and it contains a great deal of alcohol.
3. In fermentation, sugar is converted into alcohol.

Statements 1 and 2 are observations because they describe properties of the grape juice that can be tasted and smelled. Statement 3 is a conclusion because it *interprets* the observations that are available.

FIG. 1.1 A scientist working in a chemical research laboratory. Reproducible conditions in a laboratory permit experiments to yield reliable results. *(Index Stock.)*

Experimental observations lead to scientific laws

Observations we make while performing experiments are referred to as **data.** For example, if we study the behavior of gases, such as the air we breathe, we soon discover that the volume of a gas depends on a number of factors, including the mass of the gas, its temperature, and its pressure. The observations we record relating these factors are our data.

One of the goals of science is to organize facts so that relationships or generalizations among the data can be established. For instance, one generalization we would make from our observations is that when the temperature of a gas is held constant, squeezing the gas into half its original volume causes the pressure of the gas to double. If we were to repeat our experiments many times with numerous different gases, we would find that this generalization is uniformly applicable to all of them. Such a broad generalization, based on the results of many experiments, is called a **law** or **scientific law.**

We would say that the pressure of the gas is inversely proportional to its volume; the smaller the volume, the larger the pressure.

We often express laws in the form of mathematical equations. For example, if we represent the pressure of a gas by the symbol P and its volume by V, the inverse relationship between pressure and volume can be written as

$$P = \frac{C}{V}$$

where C is a proportionality constant. (We will discuss gases and the laws relating to them in greater detail in Chapter 10.)

Hypotheses and theories are models of nature

As useful as they may be, laws only state what happens; they do not provide explanations. *Why*, for example, are gases so easily compressed to a smaller volume? More specifically, *what must gases be like at the most basic, elementary level for them to behave as they do*? Answering such questions when they first arise is no simple task and requires much speculation. But over time scientists build mental pictures, called **theoretical models,** that enable them to explain observed laws.

In the development of a theoretical model, researchers form tentative explanations called **hypotheses** (Figure 1.2). They then perform experiments that test predictions derived from the model. Sometimes the results show the model is wrong. When this happens, the model must be abandoned or modified to account for the new data. Eventually, if the model

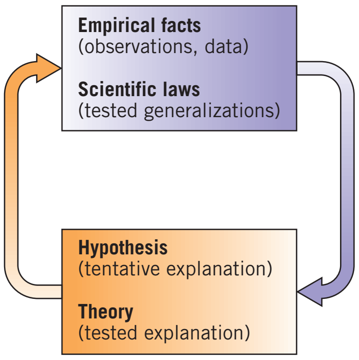


FIG. 1.2 The scientific method is cyclical. Observations suggest explanations, which suggest new experiments, which suggest new explanations, and so on.

survives repeated testing, it achieves the status of a theory. *A* **theory** *is a tested explanation of the behavior of nature.* You should keep in mind, however, that it is impossible to perform every test that might show a theory to be wrong, so we can never prove *absolutely* that a theory is correct.

Science doesn't always proceed in the orderly stepwise fashion described above. Luck sometimes plays an important role. For example, in 1828 Frederick Wöhler, a German chemist, was testing one of his theories and obtained an unexpected material when he heated a substance called ammonium cyanate. Out of curiosity he analyzed it and found it to be urea (a component of urine). This was exciting because it was the first time anyone had knowingly made a substance produced only by living creatures from a chemical not having a life origin. The fact that this could be done led to the beginning of a whole branch of chemistry called *organic chemistry*. Yet, had it not been for Wöhler's curiosity and his application of the scientific method to his unexpected results, the importance of his experiment might have gone unnoticed.

Many breakthrough discoveries in science have come about by accident.

As a final note, it is significant that the most spectacular and dramatic changes in science occur when major theories are proved to be wrong. Although this happens only rarely, when it occurs, scientists are sent scrambling to develop new theories, and exciting new frontiers are opened.

The atomic theory is a model of nature

Virtually every scientist would agree that the most significant theoretical model of nature ever formulated is the atomic theory. According to this theory, which we will discuss further in Chapter 2, all chemical substances are composed of tiny particles that we call **atoms.** Individual atoms combine in diverse ways to form more complex particles called **molecules.** Consider, for example, the substance water. Experimental evidence suggests that water molecules are each composed of two atoms of hydrogen and one of oxygen. To aid in our understanding and to help visualize how atoms combine, we often use drawings such as Figure 1.3. According to what we wish to emphasize, a variety of ways are used to describe the structures of molecules, as illustrated in Figure 1.4 for molecules of methane (the combustible fuel in natural gas).

Today we know a great deal about atoms and how they combine to form more complex materials. In coming chapters you will learn how we've come to apply this knowledge to making connections between what we physically observe in our large, *macroscopic* world and what we believe takes place in the tiny, submicroscopic world of atoms and molecules.

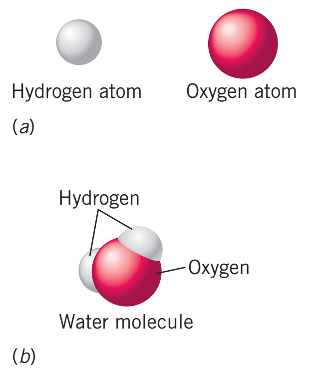


FIG. 1.3 Atoms combine to form molecules. Illustrated here is a molecule of water, which consists of one atom of oxygen and two atoms of hydrogen. (*a*) Colored spheres are used to represent individual atoms, white for hydrogen and red for oxygen. (*b*) A drawing that illustrates the shape of a water molecule.

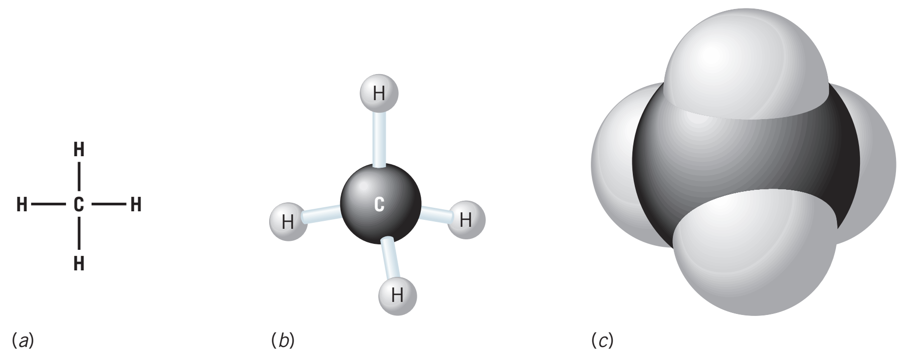


FIG. 1.4 Some of the different ways that the structures of molecules are represented. (*a*) A structure using chemical symbols to stand for atoms and dashes to indicate how the atoms are connected to each other. The molecule is methane, the substance present in natural gas that fuels stoves and Bunsen burners. A methane molecule is composed of one atom of carbon (C) and four atoms of hydrogen (H). (*b*) A *ball-and-stick model* of methane. The black ball is the carbon atom and the white balls are hydrogen atoms. (*c*) A *space-filling model* of methane that shows the relative sizes of the C and H atoms. Ball-and-stick and space-filling models are used to illustrate the three-dimensional shapes of molecules.

Learning to appreciate how chemists interpret behavior on a macroscopic level in terms of the composition of substances on an atomic scale should be one of your major goals in studying this course.

Macroscopic refers to objects large enough to be observed with the naked eye. Here we use the term to mean the things we observe with our senses, whether it be in the laboratory or in the world we encounter in our day-to-day living.

1.3 MATTER IS COMPOSED OF ELEMENTS, COMPOUNDS, AND MIXTURES

Earlier we described chemistry as being concerned with the properties and transformations of matter. **Matter** is *anything that occupies space and has mass*. It is the stuff our universe is made of, and all of the chemicals that make up tangible things, from rocks to pizza to people, are examples of matter.

In this definition, we've used the term *mass* rather than *weight*. The words mass and weight are often used interchangeably even though they refer to different things. **Mass** *refers to how much matter there is in a given object,*[2] *whereas* **weight** *refers to the force with which the object is attracted by gravity*. For example, a golf ball contains a certain amount of matter and has a certain mass, which is the same regardless of the golf ball's location. However, a golf ball on earth weighs about six times more than on the moon because the gravitational attraction of the earth is six times that of the moon. Because mass does not vary from place to place, we use mass rather than weight when we specify the amount of matter in an object. Mass is measured with an instrument called a balance, which we will discuss in Section 1.5.

Elements cannot be decomposed into simpler substances by chemical reactions

Chemistry is especially concerned with **chemical reactions,** which are *transformations that alter the chemical compositions of substances*. An important type of chemical reaction is **decomposition** in which one substance is changed into two or more others. For example, if we pass electricity through molten (melted) sodium chloride (salt), the silvery metal sodium and the pale green gas, chlorine, are formed. This change has decomposed sodium chloride into two simpler substances. No matter how we try, however, sodium and chlorine cannot be decomposed further by chemical reactions into still simpler substances that can be stored and studied.

In chemistry, *substances that cannot be decomposed into simpler materials by chemical reactions are called* **elements.** Sodium and chlorine are two examples. Others you may be familiar with include iron, aluminum, sulfur, and carbon (as in charcoal). Some elements are gases at room temperature. Examples include chlorine, oxygen, hydrogen, nitrogen, and helium. Elements are the simplest forms of matter that chemists work with directly. All more complex substances are composed of elements in various combinations.

Chemical symbols are used to identify elements

So far, scientists have discovered 90 existing elements in nature and have made 27 more, for a total of 117. Each element is assigned a unique **chemical symbol,** which can be used as an abbreviation for the name of the element. Chemical symbols are also used to stand for atoms of elements when we write *chemical formulas* such as H_2O (water) and CO_2 (carbon dioxide). We will have a lot more to say about formulas later.

In most cases, an element's chemical symbol is formed from one or two letters of its English name. For instance, the symbol for carbon is C, for bromine it is Br, and for silicon it is Si. For some elements, the symbols are derived from the non-English names given to those elements long ago. Table 1.2 contains a list of elements whose symbols come to us in that way.[3] Regardless of the origin of the symbol, the first letter is always capitalized and the second letter, if there is one, is always written lowercase. The names and chemical symbols of the elements are given on the inside front cover of the book.

[2] Mass is a measure of an object's momentum, or resistance to a change in motion. Something with a large mass, such as a truck, contains a lot of matter and is difficult to stop once it's moving. An object with less mass, such as a baseball, is much easier to stop.

[3] The symbol for tungsten is W, from the German name *wolfram*. This is the only element whose symbol is neither related to its English name nor derived from its Latin name.

TABLE 1.2 Elements That Have Symbols Derived from Their Latin Names

Element	Symbol	Latin Name	Element	Symbol	Latin Name
Sodium	Na	Natrium	Gold	Au	Aurum
Potassium	K	Kalium	Mercury	Hg	Hydrargyrum
Iron	Fe	Ferrum	Antimony	Sb	Stibium
Copper	Cu	Cuprum	Tin	Sn	Stannum
Silver	Ag	Argentum	Lead	Pb	Plumbum

Compounds are composed of two or more elements in fixed proportions

By means of chemical reactions, elements combine in various *specific proportions* to give all of the more complex substances in nature. Thus, hydrogen and oxygen combine to form water (H_2O), and sodium and chlorine combine to form sodium chloride (NaCl, common table salt). Water and sodium chloride are examples of compounds. A **compound** *is a substance formed from two or more* ***different elements*** *in which the elements are always combined in the same fixed (i.e., constant) proportions by mass.* For example, if any sample of pure water is decomposed, the mass of oxygen obtained is *always* eight times the mass of hydrogen. Similarly, when hydrogen and oxygen react to form water, the mass of oxygen consumed is always eight times the mass of hydrogen, never more and never less.

Mixtures can have variable compositions

Elements and compounds are examples of **pure substances.**[4] The composition of a pure substance is always the same, regardless of its source. Pure substances are rare, however. Usually, we encounter mixtures of compounds or elements. Unlike elements and compounds, **mixtures** *can have variable compositions.* For example, Figure 1.5 shows three mixtures that contain sugar. They have different degrees of sweetness because the amount of sugar in a given size sample varies from one to the other.

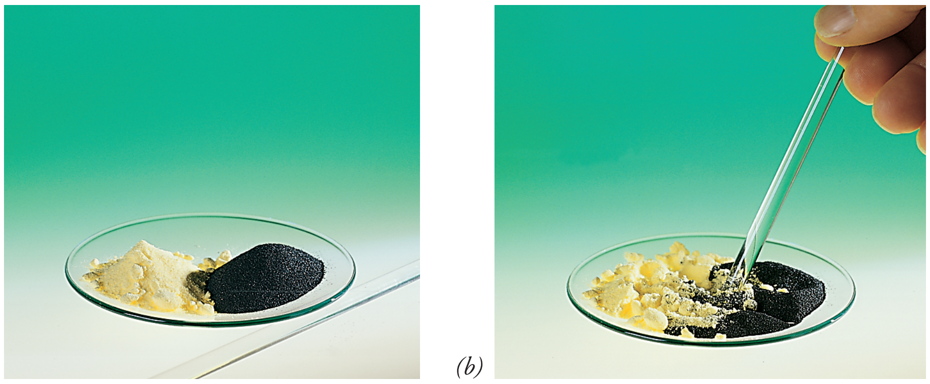
Mixtures can be either homogeneous or heterogeneous. A **homogeneous mixture** *has the same properties throughout the sample*. An example is a thoroughly stirred mixture of sugar in water. We call such a homogeneous mixture a **solution.** Solutions need not be liquids, just homogeneous. For example, the alloy used in the U.S. 5 cent coin is a solid solution of copper and nickel, and clean air is a gaseous solution of oxygen, nitrogen, and a number of other gases.

FIG. 1.5 Orange juice, Coca-Cola, and pancake syrup are mixtures that contain sugar. The amount of sugar varies from one to another because mixtures can have variable compositions. *(Thomas Brase/Stone/Getty Images; Andy Washnik; Andy Washnik.)*

[4] We have used the term *substance* rather loosely until now. Strictly speaking, **substance** really means *pure substance.* Each unique chemical element and compound is a *substance;* a mixture consists of two or more substances.

A **heterogeneous mixture** *consists of two or more regions, called* **phases,** *that differ in properties*. A mixture of olive oil and vinegar in a salad dressing, for example, is a two-phase mixture in which the oil floats on the vinegar as a separate layer (Figure 1.6). The phases in a mixture don't have to be chemically different substances like oil and vinegar, however. A mixture of ice and liquid water is a two-phase heterogeneous mixture in which the phases have the same chemical composition but occur in different *physical states* (a term we will discuss further in the next section).

FIG. 1.6 A heterogeneous mixture. The salad dressing shown here contains vinegar and vegetable oil (plus assorted other flavorings). Vinegar and oil do not dissolve in each other; instead, they form two layers. The mixture is heterogeneous because each of the separate phases (oil, vinegar, and other solids) has its own set of properties that differ from the properties of the other phases. *(Andy Washnik.)*



(*a*) (*b*)

FIG. 1.7 Formation of a mixture of iron and sulfur. (*a*) Samples of powdered sulfur and powdered iron. (*b*) A mixture of sulfur and iron is made by stirring the two powders together. *(Michael Watson.)*

The process we use to create a mixture is said to involve a **physical change,** because *no new chemical substances form*. This is illustrated in Figure 1.7 for powdered samples of the elements iron and sulfur. By simply dumping them together and stirring, the mixture forms, but both elements retain their original properties. To separate the mixture, we could similarly use just physical changes. For example, we could remove the iron by stirring the mixture with a magnet—a physical operation. The iron powder sticks to the magnet as we pull it out, leaving the sulfur behind (Figure 1.8). The mixture also could be separated by treating it with a liquid called carbon disulfide, which is able to dissolve the sulfur but not the iron. Filtering the sulfur solution from the solid iron, followed by evaporation of the liquid carbon disulfide from the sulfur solution, gives the original components, iron and sulfur, separated from each other.

The formation of a compound involves a **chemical change** (chemical reaction) because *the chemical makeup of the substances involved are changed*. Iron and sulfur, for example, combine to form a compound often called "fool's gold" because of its appearance (Figure 1.9). In this compound the elements no longer have the same properties they had before they combined, and they cannot be separated by physical means. The decomposition of fool's gold into iron and sulfur is also a chemical reaction.

The relationships among elements, compounds, and mixtures are shown in Figure 1.10.

FIG. 1.8 Formation of a mixture is a physical change. Here we see that forming the mixture has not changed the iron and sulfur into a compound of the two elements. The mixture can be separated by pulling the iron out with a magnet. *(Michael Watson.)*

FIG. 1.9 "Fool's gold." The mineral pyrite (also called iron pyrite) has an appearance that caused some miners to mistake it for real gold. *(D. Harms/Peter Arnold, Inc.)*

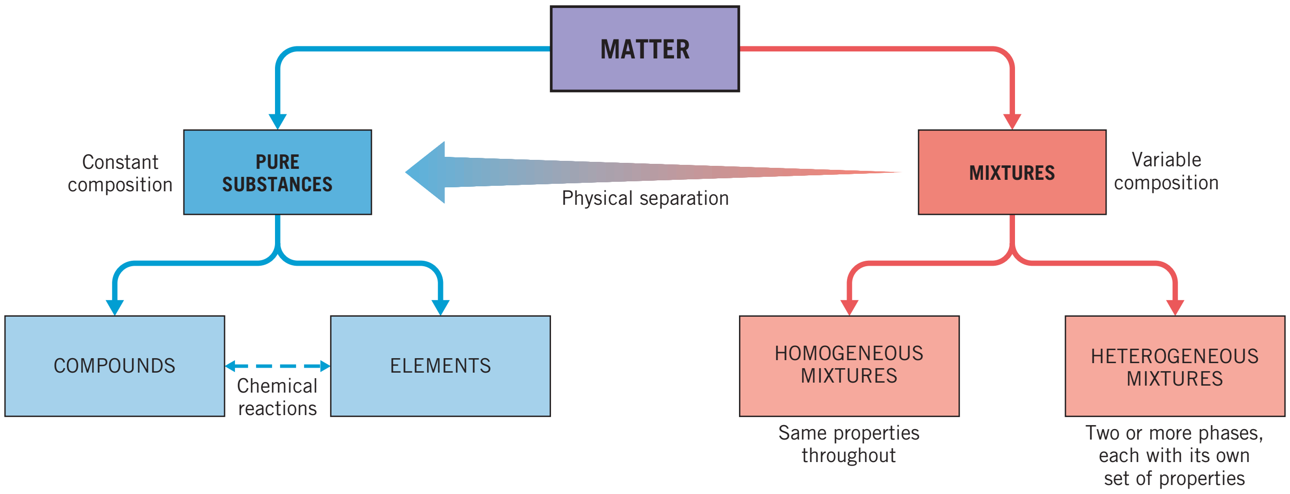


FIG. 1.10 Classification of matter.

1.4 Properties of Matter Can Be Classified in Different Ways

In chemistry we use **properties** (characteristics) of materials to identify them and to distinguish one kind from another. To help organize our thinking, we classify properties into different types.

Properties can be classified as physical or chemical

FIG. 1.11 Liquid water and ice are both composed of water molecules. Melting the ice cube doesn't change the chemical composition of the molecules. *(Susumu Sato/Corbis Images.)*

One way to classify properties is based on whether or not the chemical composition of an object is changed by the act of observing the property. *A* **physical property** *is one that can be observed without changing the chemical makeup of a substance.* For example, a physical property of gold is that it is yellow. The act of observing this property (color) doesn't change the chemical makeup of the gold. Neither does observing that gold conducts electricity, so color and electrical conductivity are physical properties.

Sometimes, observing a physical property does lead to a physical change. To measure the melting point of ice, for example, we observe the temperature at which the solid begins to melt (Figure 1.11). This is a physical change because it does not lead to a change in chemical composition; both ice and liquid water are composed of water molecules.

Solids, liquids, and gases are physical states of matter

Although ice, liquid water, and steam have quite different appearances and physical properties, they are just different forms of the same substance, water. **Solid, liquid,** and **gas** are the most common **states of matter.** As with water, most substances are able to exist in all three of these states, and the state we observe generally depends on the temperature. The obvious properties of solids, liquids, and gases can be interpreted at a submicroscopic level according to the different ways the individual atomic-size particles are organized (Figure 1.12). For a given substance, a change from one state to another is a physical change.

A chemical property describes a chemical change

A **chemical property** *describes a chemical change (chemical reaction) that a substance undergoes.* When a chemical reaction takes place, chemicals interact to form entirely *different* substances with different properties. An example is the rusting of iron, which involves a chemical reaction between iron, oxygen, and water. When the substances react, the product, rust, no longer looks like iron, oxygen, or water. It's a brown solid that isn't at all like a metal and it is not attracted by a magnet (Figure 1.13).

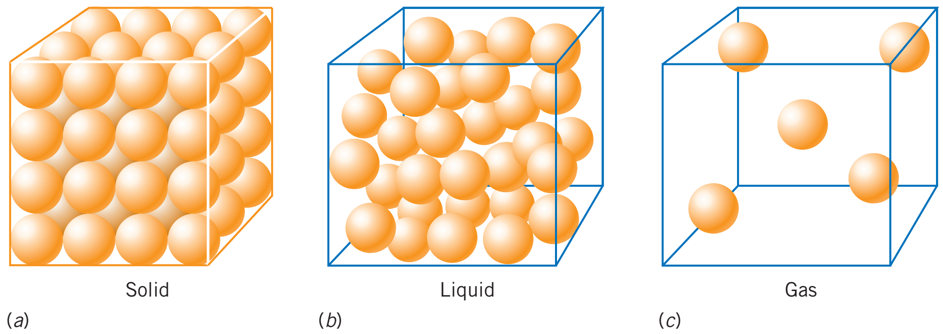


FIG. 1.12 Solid, liquid and gaseous states of matter as viewed by the atomic model of matter. (*a*) In a solid, the particles are tightly packed and cannot move easily. (*b*) In a liquid, the particles are still close together but can move past one another. (*c*) In a gas, the particles are far apart with much empty space between them.

The ability of iron to form rust in the presence of oxygen and moisture is a chemical property of iron. When we observe this property, the reaction changes the iron, oxygen, and water into rust, so after we've made the observation we no longer have the same substances as before.

FIG. 1.13 Chemical reactions cause changes in composition. Here we see a coating of rust that has formed on an iron object. The properties and chemical composition of the rust are entirely different from those of the iron. *(George B. Diebold/ Corbis Images.)*

Properties can also be classified as intensive or extensive

Another way of classifying a property is according to whether or not it depends on the size of the sample under study. For example, two different pieces of gold can have different volumes, but both have the same characteristic shiny yellow color and both will begin to melt if heated to the same temperature. Volume is said to be an **extensive property**—*a property that depends on sample size*. Color and melting point (and boiling point, too) are examples of **intensive properties**—*properties that are independent of sample size*.

Some kinds of properties are better than others for identifying substances

A job chemists often perform is *chemical analysis*. They're asked, "What is a particular sample composed of?" To answer such a question, the chemist relies on the properties of the chemicals that make up the sample. For identification purposes, intensive properties are more useful than extensive ones because every sample of a given substance exhibits the same set of intensive properties.

Color, freezing point, and boiling point are examples of intensive physical properties that can help us identify substances. Chemical properties are also intensive properties and also can be used for identification. For example, gold miners were able to distinguish between real gold and fool's gold, a mineral also called pyrite (Figure 1.9, page 7), by heating the material in a flame. Nothing happens to the gold, but the pyrite sputters, smokes, and releases bad-smelling fumes because of its ability, when heated, to react chemically with oxygen in the air.

1.5 MEASUREMENTS ARE ESSENTIAL TO DESCRIBE PROPERTIES

Observations can be qualitative or quantitative

Earlier you learned that an important step in the scientific method is observation. In general, observations fall into two categories, qualitative and quantitative. **Qualitative observations,** such as the color of a chemical or that a mixture becomes hot when a reaction occurs, do not involve numerical information and are usually of limited value. More important are **quantitative observations,** or **measurements,** which do yield numerical data. You make such observations in everyday life, for example, when you glance at your watch, or step onto a bathroom scale. In chemistry, we make various measurements that aid us in describing both chemical and physical properties.

Measurements always include units

Measurements involve numbers, but they differ from the numbers used in mathematics in two crucial ways.

First, measurements always involve a comparison. When you say that a person is six feet tall, you're really saying that the person is six times taller than a reference object that is 1 foot high, where *foot* is an example of a **unit.** Both the number and the unit are essential parts of the measurement, because the unit gives the reported value a sense of size. For example, if you were told that the distance between two points is 25, you would naturally ask "25 what?" The distance could be 25 inches, 25 feet, 25 miles, or 25 of any other unit that's used to express distance. A number without a unit is really meaningless. *Writing down a measurement without a unit is a common and serious mistake, and one you should avoid.*

The second important difference is that measurements always involve uncertainty; they are *inexact*. The act of measurement involves an estimation of one sort or another, and both the observer and the instruments used to make the measurement have inherent physical limitations. As a result, measurements always include some uncertainty, which can be minimized but never entirely eliminated. We will say more about this topic in Section 1.6.

SI units are standard in science

A standard system of units is essential if measurements are to be made consistently. In the sciences, and in every industrialized nation on earth, metric-based units are used. The advantage of working with metric units is that converting to larger or smaller units can be done simply by moving a decimal point, because metric units are related to each other by simple multiples of ten.

In 1960, a simplification of the original metric system was adopted by the General Conference on Weights and Measures (an international body). It is called the **International System of Units,** abbreviated **SI** from the French name, *Le Système International d'Unités*. The SI is now the dominant system of units in science and engineering, although there is still some usage of older metric units.

The SI has as its foundation a set of **base units** (Table 1.3) for seven measured quantities. For now, we will focus on the base units for length, mass, time, and temperature. We will discuss the unit for amount of substance, the mole, at length in Chapter 3. The unit for electrical current, the ampere, will be discussed briefly when we study electrochemistry in Chapter 19. The unit for luminous intensity, the candela, will not be important to us in this book.

Most of the base units are defined in terms of reproducible physical phenomena. For instance, the meter is defined as exactly the distance light travels in a vacuum in 1/299,792,458 of a second. Everyone has access to this standard because light and a vacuum are available to all. Only the base unit for mass is defined by an object made by human hands—a carefully preserved platinum–iridium alloy block stored at the International Bureau of Weights and Measures in France (Figure 1.14). This block serves indirectly as the calibrating standard for all "weights" used for scales and balances in the world.[5]

FIG. 1.14 The international standard kilogram. This standard for mass in the SI is made of a platinum–iridium alloy and is kept at the International Bureau of Weights and Measures in France. Other nations such as the United States maintain their own standard masses that have been carefully calibrated against this international standard. *(Courtesy Bureau International des Poids et Mesures.)*

TABLE 1.3 The SI Base Units

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

[5] Scientists are working on a method of accurately counting atoms whose masses are accurately known. Their goal is to develop a new definition of the kilogram that doesn't depend on an object that can be stolen, lost, or destroyed.

All SI units are built from the base units

The SI units for *any* physical quantity can be built from these seven base units. For example, there is no SI base unit for area, but we know that to calculate the area of a rectangular room we multiply its length by its width. Therefore, the *unit* for area is derived by multiplying the *unit* for length by the *unit* for width. Length (or width) is a base measured quantity in the SI and has the *meter* (m) as its base unit.

$$\text{length} \times \text{width} = \text{area}$$
$$(\text{meter}) \times (\text{meter}) = (\text{meter})^2$$
$$\text{m} \times \text{m} = \text{m}^2$$

The SI **derived unit** for area is therefore m^2 (read as *meters squared*, or *square meter*).

In deriving SI units, we employ a very important concept that we will use repeatedly throughout this book when we perform calculations: *Units undergo the same kinds of mathematical operations that numbers do.* We will see how this fact can be used to convert from one unit to another in Section 1.7.

EXAMPLE 1.1
Deriving SI Units

Linear momentum is a measure of the "push" a moving object has, equal to the object's mass times its velocity. What is the SI derived unit for linear momentum?

A Word about Problem Solving This is the first of many encounters you will have with solving problems in chemistry. Helping you learn how to approach and solve problems is one of the major goals of this textbook. We view problem solving as a three-step process. The first step is figuring out what has to be done to solve the problem, which is the function of the *Analysis* step described below. The second is actually performing whatever is required to obtain the answer (the *Solution* step). And finally, we examine the answer to determine whether it seems to be *reasonable*. For more information on the aids that are available to assist you in problem solving, we recommend that you read the "To the Student" section at the beginning of the book.

ANALYSIS: To derive a unit for a quantity we must first express it in terms of simpler quantities. We're told that linear momentum is mass times velocity. Therefore, the SI unit for linear momentum will be the SI unit for mass times the SI unit for velocity. The SI unit for mass is the kilogram (kg). Velocity is distance traveled (length) per unit time, so it has derived SI units of meters per second, m/s. Multiplying these units should give the derived unit for linear momentum.

SOLUTION:

$$\text{mass} \times \text{velocity} = \text{linear momentum}$$
$$\text{mass} \times \text{length/time} = \text{linear momentum}$$
$$\text{kilogram} \times \text{meter/second} = \text{kilogram meter/second}$$
$$\text{kg} \times \text{m/s} = \text{kg m/s}$$

IS THE ANSWER REASONABLE? *Before leaving a problem, it is always wise to examine the answer to see whether it makes sense.* For numerical calculations, ask yourself, "Is the answer too large, or too small?" Judging the answers to such questions serves as a check on the arithmetic as well as on the method of obtaining the answer and can help you find obvious errors. In this problem, the check is simple. The derived unit for linear momentum should be the product of units for mass and velocity, and this is obviously true. Therefore, our answer is correct.

Practice Exercise 1:[6] The volume of a sphere is given by the formula $V = \frac{4}{3}\pi r^3$, where r is the radius of the sphere. From this equation, determine the SI unit for volume. (Hint: r is a distance, so it must have a distance unit.)

[6] Answers to the Practice Exercises are found at the back of the book.

Practice Exercise 2: When you "step hard on the gas" in a car you feel an invisible force pushing you back in your seat. This force equals the product of your mass, m, times the acceleration, a, of the car. In equation form, this is $F = ma$. Acceleration is the change in velocity, v, with time, t:

$$a = \frac{\text{change in } v}{\text{change in } t}$$

Therefore, the units of acceleration are those of velocity divided by time. Velocity is a ratio of distance divided by time, $v = d/t$, so the units of velocity are those of distance divided by time. What is the SI derived unit for force expressed in SI base units?

We can construct SI units of any convenient size using decimal multipliers

Sometimes the basic units are either too large or too small to be used conveniently. For example, the meter is inconvenient for expressing the size of very small things such as bacteria. The SI solves this problem by forming larger or smaller units by applying **decimal multipliers** to the base units. Table 1.4 lists the most commonly used decimal multipliers and the prefixes used to identify them.

When the name of a unit is preceded by one of these prefixes, the size of the unit is modified by the corresponding decimal multiplier. For instance, the prefix *kilo-* indicates a multiplying factor of 10^3, or 1000. Therefore, a *kilo*meter is a unit of length equal to 1000 meters.[7] The symbol for kilometer (km) is formed by applying the symbol meaning kilo (k) as a prefix to the symbol for meter (m). Thus 1 km = 1000 m (or alternatively, 1 km = 10^3 m). Similarly a decimeter (dm) is 1/10 of a meter, so 1 dm = 0.1 m (1 dm = 10^{-1} m).

The symbols and multipliers listed in colored, boldface type in Table 1.4 are the ones most commonly encountered in chemistry.

TOOLS
SI prefixes

TABLE 1.4 Decimal Multipliers That Serve as SI Prefixes[a]

Prefix	Meaning	Symbol	Multiplication factor (fraction)	Multiplication factor (power of ten)
exa		E		10^{18}
peta		P		10^{15}
tera		T		10^{12}
giga		G		10^{9}
mega	**millions of**	**M**	**1,000,000**	$\mathbf{10^{6}}$
kilo	**thousands of**	**k**	**1000**	$\mathbf{10^{3}}$
hecto		h		10^{2}
deka		da		10^{1}
deci	**tenths of**	**d**	**0.1**	$\mathbf{10^{-1}}$
centi	**hundredths of**	**c**	**0.01**	$\mathbf{10^{-2}}$
milli	**thousandths of**	**m**	**0.001**	$\mathbf{10^{-3}}$
micro	**millionths of**	**μ**	**0.000001**	$\mathbf{10^{-6}}$
nano	**billionths of**	**n**	**0.000000001**	$\mathbf{10^{-9}}$
pico	**trillionths of**	**p**	**0.000000000001**	$\mathbf{10^{-12}}$
femto		f		10^{-15}
atto		a		10^{-18}

[a]Be sure you learn the prefixes shown in bold colored type.

[7] In the sciences, powers of 10 are often used to express large and small numbers. The quantity 10^3 means $10 \times 10 \times 10 = 1000$. Similarly, the quantity $6.5 \times 10^2 = 6.5 \times 100 = 650$. Numbers less than 1 have negative exponents when expressed as powers of 10. Thus, the fraction $\frac{1}{10}$ is expressed as 10^{-1}, so the quantity 10^{-3} means $\frac{1}{10} \times \frac{1}{10} \times \frac{1}{10} = \frac{1}{1000} = 0.001$. A value of $6.5 \times 10^{-3} = 6.5 \times 0.001 = 0.0065$. Numbers written as 6.5×10^2 and 6.5×10^{-3}, with the decimal point between the first and second digit, are said to be expressed in **standard scientific notation.**

Non-SI units are still in common use

Some older metric units that are not part of the SI system are still used in the laboratory and in the scientific literature. Some of these units are listed in Table 1.5; others will be introduced as needed in upcoming chapters.

The United States is the only large nation still using the *English system* of units, which measures distance in inches, feet, and miles; volume in ounces, quarts, and gallons; and mass in ounces and pounds. However, a gradual transition to metric units is occurring. Beverages, food packages, tools, and machine parts are often labeled in both English and metric units (Figure 1.15). Common conversions between the English system and the SI are given in Table 1.6 and inside the rear cover of the book.[8]

FIG. 1.15 Metric units are becoming commonplace on many consumer products. *(Michael Watson.)*

TABLE 1.5 Some Non-SI Metric Units Commonly Used in Chemistry

Measurement	Name	Symbol	Value in SI units
Length	angstrom	Å	1 Å = 0.1 nm = 10^{-10} m
Mass	atomic mass unit	u (amu)	1 u = 1.66054×10^{-27} kg, approximately
	metric ton	t	1 t = 10^3 kg
Time	minute	min	1 min = 60 s
	hour	h (hr)	1 h = 60 min = 3600 s
Temperature	degree Celsius	°C	Add 273.15 to obtain the Kelvin temperature
Volume	liter	L	1 L = 1000 cm^3

TABLE 1.6 Some Useful Conversions

Measurement	English to Metric	Metric to English
Length	1 in. = 2.54 cm	1 m = 39.37 in.
	1 yd = 0.9144 m	1 km = 0.6215 mi
	1 mi = 1.609 km	
Mass	1 lb = 453.6 g	1 kg = 2.205 lb
	1 oz = 28.35 g	
Volume	1 gal = 3.785 L	1 L = 1.057 qt
	1 qt = 946.4 mL	
	1 oz (fluid) = 29.6 mL	

We use several common units in laboratory measurements

TOOLS *Units for laboratory measurements*

The most common measurements you will make in the laboratory will be those of length, volume, mass, and temperature.

Length

The SI base unit for length, the **meter (m),** is too large for most laboratory purposes. More convenient units are the **centimeter (cm)** and the **millimeter (mm).** They are related to the meter as follows.

$$1 \text{ cm} = 10^{-2} \text{ m} = 0.01 \text{ m}$$

$$1 \text{ mm} = 10^{-3} \text{ m} = 0.001 \text{ m}$$

It is also useful to know the relationships

$$1 \text{ m} = 100 \text{ cm} = 1000 \text{ mm}$$

$$1 \text{ cm} = 10 \text{ mm}$$

An older non-SI unit called the angstrom (Å) is often used to describe the dimensions of atomic and molecular sized particles: 1 Å = 0.1 nm = 10^{-10} m

[8] Originally, these conversions were established by measurement. For example, if a metric ruler is used to measure the length of an inch, it is found that 1 in. equals 2.54 cm. Later, to avoid confusion about the accuracy of such measurements, it was agreed that these relationships would be taken to be exact. For instance, 1 in. is now defined as *exactly* 2.54 cm. Exact relationships also exist for the other quantities, but for simplicity many have been rounded off. For example, 1 lb = 453.59237 g, *exactly*.

FIG. 1.16 Common laboratory glassware used for measuring volumes. Graduated cylinders are used to measure volumes to the nearest milliliter. Precise measurements of volume are made using burets, pipets, and volumetric flasks. *(Andy Washnik.)*

Volume

Volume is a derived unit with dimensions of (length)^3. With these dimensions expressed in meters, the derived SI unit for volume is the **cubic meter, m^3.**

In chemistry, measurements of volume usually arise when we measure amounts of liquids. The traditional metric unit of volume used for this is the **liter (L).** In SI terms, a liter is defined as exactly 1 cubic decimeter.

$$1 \text{ L} = 1 \text{ dm}^3$$

However, even the liter is too large to conveniently express most volumes measured in the lab. The glassware we normally use, such as that illustrated in Figure 1.16, is marked in **milliliters (mL).**[9]

$$1 \text{ L} = 1000 \text{ mL}$$

Because 1 dm = 10 cm, then 1 dm^3 = 1000 cm^3. Therefore, 1 mL is exactly the same as 1 cm^3.

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

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

Sometimes you may see cm^3 abbreviated cc (especially in medical applications), although the SI frowns on this symbol. Figure 1.17 compares the cubic meter, liter, and milliliter.

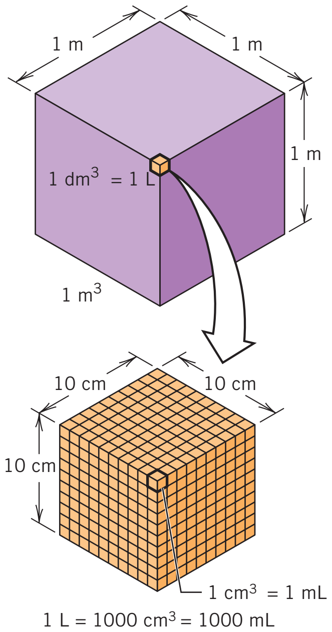


FIG. 1.17 Comparing volume units. A cubic meter (m^3) is approximately equal to a cubic yard, 1000 cm^3 is approximately a quart, and 1 cm^3 is approximately $\frac{1}{30}$ of a fluid ounce.

Mass

In the SI, the base unit for mass is the **kilogram (kg),** although the **gram (g)** is a more conveniently sized unit for most laboratory measurements. One gram, of course, is $\frac{1}{1000}$ of a kilogram (1 kilogram = 1000 g, so 1 g must equal 0.001 kg).

Mass is measured by comparing the weight of a sample with the weights of known standard masses. The operation is called **weighing,** and the apparatus used is called a **balance** (Figure 1.18). For the balance in Figure 1.18*a*, we would place our sample on the left pan and then add standard masses to the other. When the weight of the sample and the total weight of the standards are in balance (when they match), their masses are then equal. Figure 1.19 gives the masses of some common objects in SI units.

[9] Use of the abbreviations L for liter and mL for milliliter is rather recent. Confusion between the printed letter 1 and the number 1 prompted the change from l for liter and ml for milliliter. You may encounter the abbreviation ml in other books or on older laboratory glassware.

(a) *(b)* *(c)*

FIG. 1.18 **Typical laboratory balances.** (*a*) A traditional two-pan analytical balance capable of measurements to the nearest 0.0001 g. (*b*) A modern top-loading balance capable of mass measurements to the nearest 0.001 g (fitted with a cover to reduce the effects of air currents and thereby improve precision). (*c*) A modern analytical balance capable of measurements to the nearest 0.0001 g. *(Michael Watson; Courtesy Central Scientific Co.; Courtesy Cole-Parmer Instrument Co.)*

FIG. 1.19 **Masses of several common objects in metric and English units.** *(Coco McCoy/Rainbow; Coco McCoy/Rainbow; Andy Washnik; Jim Cummins/Taxi/Getty Images.)*

Temperature

Temperature is usually measured with a thermometer (Figure 1.20). Thermometers are graduated in *degrees* according to one of two temperature scales. Both scales use as reference

(a)

(b)

FIG. 1.20 **Typical laboratory thermometers.** (*a*) A traditional mercury thermometer. (*b*) An electronic thermometer. *(Michael Watson/Corbis Images.)*

In chemistry, reference data are commonly tabulated at 25 °C, which is close to room temperature. Biologists often carry out their experiments at 37 °C because that is normal human body temperature.

points the temperature at which water freezes[10] and the temperature at which it boils. On the **Fahrenheit scale** water freezes at 32 °F and boils at 212 °F. If you've been raised in the United States, this is probably the scale you're most familiar with. In recent times, however, you have probably noticed an increased use of the Celsius scale, especially in weather broadcasts. This is the scale we use most often in the sciences. On the **Celsius scale** water freezes at 0 °C and boils at 100 °C. (See Figure 1.21.)

As you can see in Figure 1.21, on the Celsius scale there are 100 degree units between the freezing and boiling points of water, while on the Fahrenheit scale this same temperature range is spanned by 180 degree units. Consequently, 5 Celsius degrees are the same as 9 Fahrenheit degrees. We can use the following equation as a tool to convert between these temperature scales.

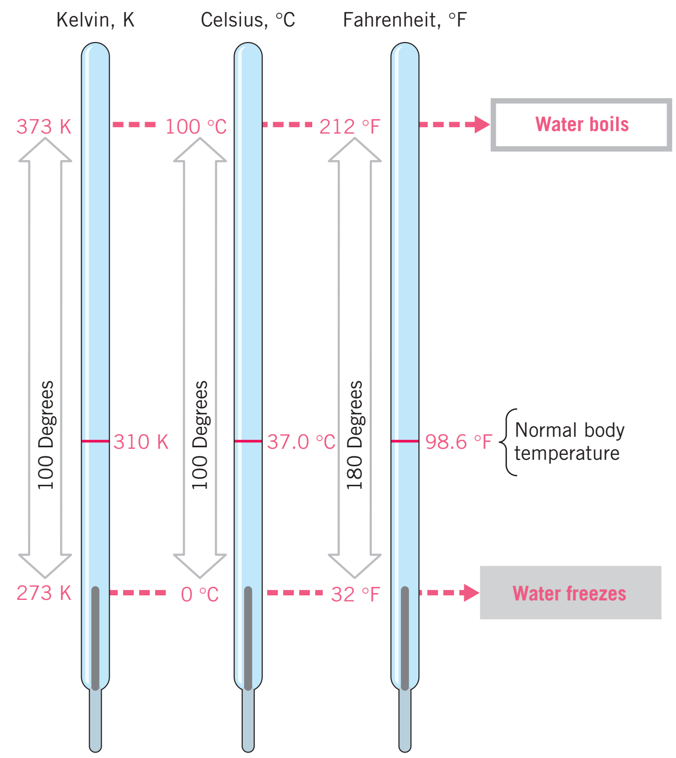
TOOLS
Celsius to Fahrenheit conversions

$$t_F = \left(\frac{9\ °\text{F}}{5\ °\text{C}}\right)t_C + 32\ °\text{F} \tag{1.1}$$

In this equation, t_F is the Fahrenheit temperature and t_C is the Celsius temperature. As noted earlier, units behave like numbers in calculations, and we see in Equation 1.1 that °C "cancels out" to leave only °F. The 32 °F is added to account for the fact that the freezing point of water (0 °C) occurs at 32 °F on the Fahrenheit scale. Equation 1.1 can easily be rearranged to permit calculating °C from °F.

We will use a capital T to stand for the Kelvin temperature and a lowercase t (as in t_C) to stand for the Celsius temperature. This conforms to usage described by the International Bureau of Weights and Measures in Sèvres, France, and the National Institute of Standards and Technology (NIST) in Gaithersburg, MD.

The SI unit of temperature is the **kelvin (K),** which is the degree unit on the **Kelvin temperature scale.** Notice that the temperature unit is K, not °K (the degree symbol, °, is omitted). Also notice that the name of the unit, kelvin, is not capitalized. Equations that include temperature as a variable sometimes take on a simpler form when Kelvin temperatures are used. We will encounter this situation many times throughout the book.



The name of the temperature scale, the Kelvin scale, is capitalized, but the name of the unit, the kelvin, is not. However, the symbol for the kelvin is the capital letter K.

FIG. 1.21 Comparison among Kelvin, Celsius, and Fahrenheit temperature scales.

[10] Water freezes and ice melts at the same temperature, and a mixture of ice and liquid water will maintain a constant temperature of 32 °F or 0 °C. If heat is added, some ice melts; if heat is removed, some liquid water freezes, but the temperature doesn't change. This constancy of temperature is what makes the "ice point" convenient for calibrating thermometers.

Figure 1.21 shows how the Kelvin, Celsius, and Fahrenheit temperature scales relate to each other. Notice that the kelvin is *exactly* the same size as the Celsius degree. *The only difference between these two temperature scales is the zero point.* The zero point on the Kelvin scale is called **absolute zero** and corresponds to nature's coldest temperature. It is 273.15 degree units below the zero point on the Celsius scale, which means that 0 °C equals 273.15 K, and 0 K equals −273.15 °C. Thermometers are never marked with the Kelvin scale, so to convert from Celsius to Kelvin temperatures the following equation applies.

$$T_K = (t_C + 273.15\ ^\circ\mathrm{C})\left(\frac{1\ \mathrm{K}}{1\ ^\circ\mathrm{C}}\right) \tag{1.2}$$

TOOLS
Celsius to Kelvin conversions

This amounts to simply adding 273.15 to the Celsius temperature to obtain the Kelvin temperature. Often we are given Celsius temperatures rounded to the nearest degree, in which case we round 273.15 to 273. Thus, 25 °C equals (25 + 273) K or 298 K.

EXAMPLE 1.2
Converting among Temperature Scales

Thermal pollution, the release of large amounts of heat into rivers and other bodies of water, is a serious problem near power plants and can affect the survival of some species of fish. For example, trout will die if the temperature of the water rises above approximately 25 °C. (a) What is this temperature in °F? (b) Rounded to the nearest whole degree unit, what is this temperature in kelvins?

ANALYSIS: *Usually, the first job in solving a problem is determining which tools are required to do the work.* Both parts of the problem here deal with temperature conversions. Therefore, we ask ourselves, "What tools do we have that relate temperature scales to each other?" Let's write them:

Equation 1.1 $$t_F = \left(\frac{9\ ^\circ\mathrm{F}}{5\ ^\circ\mathrm{C}}\right) t_C + 32\ ^\circ\mathrm{F}$$

Equation 1.2 $$T_K = (t_C + 273.15\ ^\circ\mathrm{C})\left(\frac{1\ \mathrm{K}}{1\ ^\circ\mathrm{C}}\right)$$

Equation 1.1 relates Fahrenheit temperatures to Celsius temperatures, so this is the tool we need to answer part (a). Equation 1.2 relates Kelvin temperatures to Celsius temperatures, and this is the tool we need for part (b). Now that we have what we need, the rest follows.

SOLUTION:

(a) We substitute the value of the Celsius temperature (25 °C) for t_C.

$$\begin{aligned} t_F &= \left(\frac{9\ ^\circ\mathrm{F}}{5\ \cancel{^\circ\mathrm{C}}}\right)(25\ \cancel{^\circ\mathrm{C}}) + 32\ ^\circ\mathrm{F} \\ &= 77\ ^\circ\mathrm{F} \end{aligned}$$

Therefore, 25 °C = 77 °F. (Notice that we have canceled the unit °C in the equation above. As noted earlier, units behave the same as numbers do in calculations.)

(b) Once again, we have a simple substitution. Since t_C = 25 °C, the Kelvin temperature (rounded) is

$$\begin{aligned} T_K &= (25\ ^\circ\mathrm{C} + 273\ ^\circ\mathrm{C})\left(\frac{1\ \mathrm{K}}{1\ ^\circ\mathrm{C}}\right) \\ &= 298\ \cancel{^\circ\mathrm{C}}\left(\frac{1\ \mathrm{K}}{1\ \cancel{^\circ\mathrm{C}}}\right) = 298\ \mathrm{K} \end{aligned}$$

Thus, 25 °C = 298 K.

ARE THE ANSWERS REASONABLE? For part (a), we know that a Fahrenheit degree is about half the size of a Celsius degree, so 25 Celsius degrees should be about 50 Fahrenheit degrees. The positive value for the Celsius temperature tells us we have a temperature *above* the freezing point of water. Since water freezes at 32 °F, the Fahrenheit temperature should be approximately 32 °F + 50 °F = 82 °F. The answer of 77 °F is quite close.

For part (b), we recall that 0 °C = 273 K. A temperature above 0 °C must be higher than 273 K. Our calculation, therefore, appears to be correct.

Practice Exercise 3: What Fahrenheit temperature corresponds to a Celsius temperature of 86 °C? (Hint: What tool relates the two temperature scales?)

Practice Exercise 4: What Celsius temperature corresponds to 50 °F? What Kelvin temperature corresponds to 68 °F (expressed to the nearest whole kelvin unit)?

1.6 Measurements Always Contain Some Uncertainty

We noted in the preceding section that measurements are inexact; they contain **uncertainties** (also called **errors**). One source of uncertainty is associated with limitations in our ability to read the scale of the measuring instrument. Uncontrollably changing conditions at the time of the measurement can also cause errors that are more important than scale reading errors. For example, if you are measuring a length of wire with a ruler, you may not be holding the wire perfectly straight every time.

If we were to take an enormous number of measurements using appropriately adjusted instruments, statistically half of the measurements should be larger and half smaller than the true value of the measured quantity. And, in fact, we do observe that a series of measurements tend to cluster around some central value, which we generally assume is close to the true value. We can estimate the central value quite simply by reporting the **average,** or **mean,** of the series of measurements. This is done by summing the measurements and then dividing by the number of measurements we made. Although making repeated measurements is tedious, the more measurements we make, the more confident we can be that the average is close to the true value that all measurements would be grouped around.

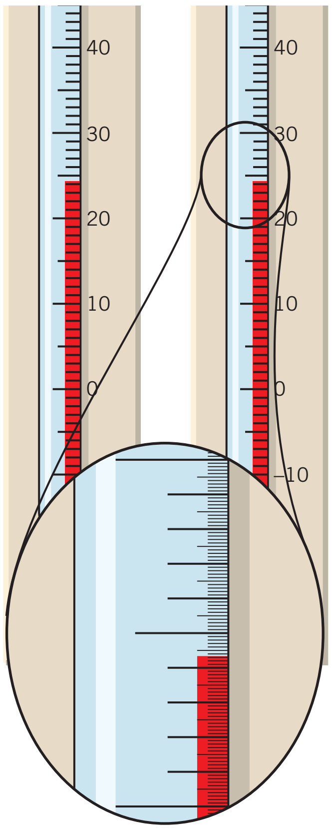


FIG. 1.22 Thermometers with different scales give readings with different precision. The thermometer on the left has marks that are one degree apart, allowing the temperature to be estimated to the nearest tenth of a degree. The thermometer on the right has marks every 0.1 °C. This scale permits estimation of the hundredths place.

Uncertainties in measurements are a natural part of reading a scale

One kind of error that can't be eliminated arises when we attempt to obtain a measurement by reading the scale on an instrument. Consider, for example, reading the same temperature from each of the two thermometers in Figure 1.22.

The marks on the left thermometer are one degree apart, and we can see that the temperature lies between 24 °C and 25 °C. When reading a scale, we always record the last digit to the nearest tenth of the smallest scale division. Looking closely, therefore, we might estimate that the fluid column falls about 3/10 of the way between the marks for 24 and 25 degrees, so we can report the temperature to be 24.3 °C. However, it would be foolish to say that the temperature is *exactly* 24.3 °C. The last digit is only an estimate, and the left thermometer might be read as 24.2 °C by one observer or 24.4 °C by another. Because different observers might obtain values that differ by 0.1 °C, there is an uncertainty of ±0.1 °C in the measured temperature. We can express this, if we wish, by writing the temperature as 24.3 ± 0.1 °C.

The thermometer on the right has marks that are 1/10 of a degree apart, which allows us to estimate the temperature as 24.32 °C. In this case, we are estimating the hundredths place and the uncertainty is ±0.01 °C. We could write the temperature as 24.32 ± 0.01 °C. Notice that because the thermometer on the right is more finely graduated, we are able to obtain measurements with smaller uncertainties. We would have more confidence in temperatures read from the thermometer on the right in Figure 1.22 because it has more digits and a smaller amount of uncertainty. *The reliability of a piece of data is indicated by the number of digits used to represent it.*

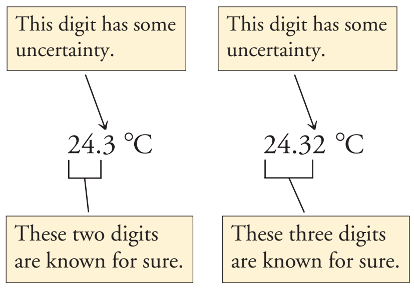
By convention in science, *all digits in a measurement up to and including the first estimated digit are recorded.* If a reading measured with the thermometer on the right seemed exactly on the 24 °C mark, we would record the temperature as 24.00 °C, not 24 °C, to show that the thermometer can be read to the nearest 1/100 of a degree.

Measurements are written using the significant figures convention

The concepts discussed above are so important that we have special terminology to describe numbers that come from measurements.

> *Digits that result from measurement such that only the digit farthest to the right is not known with certainty are called* **significant figures** (*or* **significant digits**).

The number of significant figures in a measurement is equal to the number of digits known for sure *plus* one that is estimated. Let's look at our two temperature measurements.



The first measurement, 24.3 °C, has three significant figures; the second, 24.32 °C, has four significant figures.

Accuracy is correctness; precision is reproducibility

Two words often used in reference to measurements are *accuracy* and *precision*. **Accuracy** refers to how close a measurement is to the true or correct value. **Precision** refers to how closely repeated measurements of a quantity come to each other and to the average. Notice that the two terms are not equivalent, because the average doesn't always correspond to the true or correct value. A practical example of how accuracy and precision differ is illustrated in Figure 1.23.

For measurements to be **accurate,** the measuring device must be carefully calibrated (adjusted) so it gives correct values when a standard reference is used with it. For example, to calibrate an electronic balance, a known reference mass is placed on the balance and a calibration routine within the balance is initiated. Once calibrated, the balance will give correct readings, the accuracy of which is determined by the accuracy of the standard mass used. Standard reference masses (also called "weights") can be purchased from scientific supply companies.

Precision refers to how closely repeated measurements of the same quantity come to each other. In general, the smaller the uncertainty (i.e., the "plus or minus" part of the measurement), the more precise the measurement. This translates as: *the more significant figures in a measured quantity, the more precise the measurement.*

We usually assume that a very precise measurement is also of high accuracy. We can be wrong, however, if our instruments are improperly calibrated. For example, the improperly marked ruler in Figure 1.24 might yield measurements which vary by a hundredth of a centimeter (±0.01 cm), but all the measurements would be too large by 1 cm—a case of good precision but poor accuracy.

When are zeros significant digits?

Usually, it is simple to determine the number of significant figures in a measurement; we just count the digits. Thus 3.25 has three significant figures and 56.205 has five of them. When zeros come at the beginning or the end of a number, however, they sometimes cause confusion.

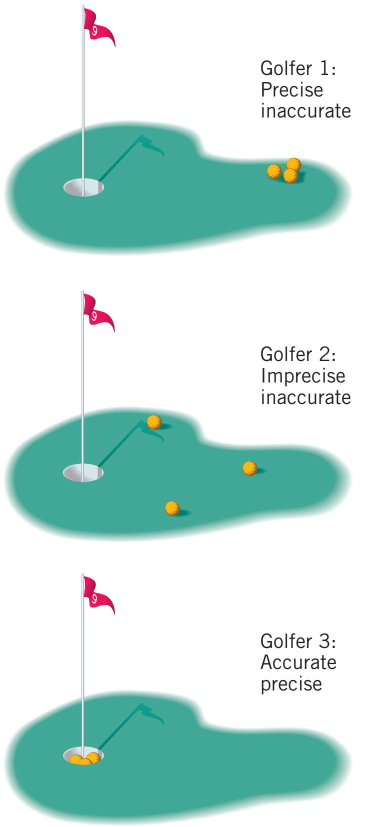
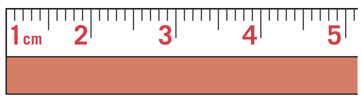


FIG. 1.23 The difference between precision and accuracy in the game of golf. Golfer 1 hits shots that are precise (because they are tightly grouped) but the accuracy is poor because the balls are not near the target (the "true" value). Golfer 2 needs help. His shots are neither precise nor accurate. Golfer 3 wins the prize with shots that are precise (tightly grouped) and accurate (in the hole).



How accurate would measurements be with this ruler?

FIG. 1.24 An improperly marked ruler. This improperly marked ruler will yield measurements that are each wrong by one whole unit. The measurements might be precise, but the accuracy would be very poor.

TOOLS
Counting significant figures

Zeros to the right of a decimal point are always counted as significant. Thus, 4.500 m has four significant figures because the zeros would not be written unless those digits were known to be zeros.

Zeros to the left of the first nonzero digit are never counted as significant. For instance, a length of 2.3 mm is the same as 0.0023 m. Since we are dealing with the same measured value, its number of significant figures cannot change when we change the units. Both quantities have two significant figures.

Zeros on the end of a number without a decimal point are assumed not to be significant. For example, suppose you were told that a protest march was attended by 45,000 people. If this was just a rough estimate, it might be uncertain by as much as several thousand, in which case the value 45,000 represents just two significant figures, since the "5" is the uncertain digit. None of the zeros would then count as significant figures. On the other hand, suppose the protesters were carefully counted using an aerial photograph, so that the count could be reported to be 45,000 give or take about 100 people. In this case, the value represents 45,000 ± 100 protesters and contains three significant figures, with the uncertain digit being the zero in the hundreds place. So a simple statement such as "there were 45,000 people attending the march" is ambiguous. We can't tell how many significant digits the number has from the number alone. We can be sure that the nonzero digits are significant, though. The best we can do is to say, "45,000 has *at least* two significant figures."

We can avoid confusion by using *scientific notation* when we report a measurement. For example, if we want to report the number of protesters as 45,000 give or take a thousand, we can write the rough estimate as 4.5×10^4. The 4.5 shows the number of significant figures and the 10^4 tells us the location of the decimal. The value obtained from the aerial photograph count, on the other hand, can be expressed as 4.50×10^4. This time the 4.50 shows three significant figures and an uncertainty of ±100 people.

Measurements limit the precision of results calculated from them

When several measurements are obtained in an experiment they are usually combined in some way to calculate a desired quantity. For example, to determine the area of a rectangular carpet we require two measurements, length and width, which are then multiplied to give the answer we want. To get some idea of how precise the area really is, we need a way to take into account the precision of the various values used in the calculation. To make sure this happens, we follow certain rules according to the kinds of arithmetic being performed.

Multiplication and division

TOOLS

Significant figures: multiplication and division

For multiplication and division, the number of significant figures in the answer should not be greater than the number of significant figures in the least precise measurement. Let's look at a typical problem involving some measured quantities.

$$\frac{\overset{\boxed{\text{3 sig. figures}}}{3.14} \times \overset{\boxed{\text{4 sig. figures}}}{2.751}}{\underset{\boxed{\text{2 sig. figures}}}{0.64}} = 13$$

The result displayed on a calculator[11] is 13.49709375. However, the least precise factor, 0.64, has only two significant figures, so the answer should have only two. The correct answer, 13, is obtained by rounding off the calculator answer.[12]

[11] Calculators usually give too many significant figures. An exception is when the answer has zeros at the right that are significant figures. For example, an answer of 1.200 would be displayed on most calculators as 1.2. If the zeros belong in the answer, be sure to write them down.

[12] When we wish to round off a number at a certain point, we simply drop the digits that follow if the first of them is less than 5. Thus, 8.1634 rounds to 8.16, if we wish to have only two decimal places. If the first digit after the point of round off is larger than 5, or if it is 5 followed by other nonzero digits, then we add 1 to the preceding digit. Thus 8.167 and 8.1653 both round to 8.17. Finally, when the digit after the point of round off is a 5 and no other digits follow the 5, then we drop the 5 if the preceding digit is even and add 1 if it is odd. Thus, 8.165 rounds to 8.16 and 8.175 rounds to 8.18.

Addition and subtraction

TOOLS Significant figures: addition and subtraction

For addition and subtraction, the answer should have the same number of decimal places as the quantity with the fewest number of decimal places. As an example, consider the following addition of measured quantities.

$$\begin{array}{r} 3.247 \\ 41.36 \\ +\ 125.2 \\ \hline 169.8 \end{array}$$

+ 125.2 ⟵ (This number has only 1 decimal place.)

169.8 ⟵ (The answer has been rounded to 1 decimal place.)

In this calculation, the digits beneath the 6 and the 7 are unknown; they could be anything. (They're not necessarily zeros because if we *knew* they were zeros, then zeros would have been written there.) Adding an unknown digit to the 6 or 7 will give an answer that's also unknown, so for this sum we are not justified in writing digits in the second and third places after the decimal point. Therefore, we round the answer to the nearest tenth.

Exact numbers contain no uncertainty

TOOLS Significant figures: exact numbers

Numbers that come from definitions, such as 12 in. = 1 ft, and those that come from a direct count, such as the number of people in a room, have no uncertainty. We can assume that such **exact numbers** have an infinite number of significant figures. Therefore, we ignore exact numbers when applying the rules described above.

Practice Exercise 5: Perform the following calculations involving measurements and round the results so they have the correct number of significant figures and proper units. (Hint: Apply the rules for significant figures described in this section and keep in mind that units behave like numbers do in calculations.)

(a) 21.0233 g + 21.0 g

(b) 10.0324 g/11.7 mL

(c) $\dfrac{14.25 \text{ cm} \times 12.334 \text{ cm}}{(2.223 \text{ cm} - 1.04 \text{ cm})}$

Practice Exercise 6: Perform the following calculations involving measurements and round the results so that they are written to the correct number of significant figures and have the correct units.

(a) 32.02 mL − 2.0 mL

(b) 54.183 g − 0.0278 g

(c) 10.0 g + 1.03 g + 0.243 g

(d) $43.4 \text{ in.} \times \dfrac{1 \text{ ft}}{12 \text{ in.}}$

(e) $\dfrac{1.03 \text{ m} \times 2.074 \text{ m} \times 3.9 \text{ m}}{12.46 \text{ m} + 4.778 \text{ m}}$

1.7 UNITS CAN BE CONVERTED USING THE FACTOR-LABEL METHOD

After analyzing a problem and assembling the necessary information to solve it, the next step is working the problem to obtain an answer. For numerical problems, scientists usually use a system called the **factor-label method** (also called **dimensional analysis**) to help them perform the correct arithmetic. As you will see, often this method also helps in analyzing the problem and selecting the tools needed to solve it.

In the factor-label method we treat a numerical problem as one involving a conversion of units from one kind to another. To do this we use one or more *conversion factors* to change the units of the given quantity to the units of the answer.

$$(\text{Given quantity}) \times (\text{conversion factor}) = (\text{desired quantity})$$

A **conversion factor** *is a fraction formed from a valid relationship or equality between units and is used to switch from one system of measurement and units to another.* To illustrate, suppose we want to express a person's height of 72.0 inches in centimeters. To do this we need the relationship between the inch and the centimeter. We can obtain this from Table 1.6.

$$2.54 \text{ cm} = 1 \text{ in. (exactly)} \tag{1.3}$$

To construct a valid conversion factor, the relationship between the units must be true. For example, the statement

$$3 \text{ ft} = 41 \text{ in.}$$

is false. Although you might make a conversion factor out of it, any answers you would calculate are sure to be incorrect. *Correct answers require correct relationships between units.*

If we divide both sides of this equation by 1 in., we obtain a conversion factor.

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

Notice that we have canceled the units from both the numerator and denominator of the center fraction, leaving the first fraction equaling 1. As mentioned earlier, *units behave just as numbers do in mathematical operations;* this is a key part of the factor-label method. Let's see what happens if we multiply 72.0 inches, the height that we mentioned, by this fraction.

The relationship between the inch and the centimeter is exact, so the numbers in 1 in. = 2.54 cm have an infinite number of significant figures.

$$72.0\ \cancel{\text{in.}} \times \frac{2.54\ \text{cm}}{1\ \cancel{\text{in.}}} = 183\ \text{cm}$$

$$\left(\begin{array}{c}\text{given}\\ \text{quantity}\end{array}\right) \times \left(\begin{array}{c}\text{conversion}\\ \text{factor}\end{array}\right) = \left(\begin{array}{c}\text{desired}\\ \text{quantity}\end{array}\right)$$

Because we have multiplied 72.0 in. by something that is equal to 1, we know we haven't changed the magnitude of the person's height. We have, however, changed the units. Notice that we have canceled the unit inches. The only unit left is centimeters, which is the unit we want for the answer. The result, therefore, is the person's height in centimeters.

One of the benefits of the factor-label method is that it often lets you know when you have done the *wrong* arithmetic. From the relationship in Equation 1.3, we can actually construct two conversion factors:

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

We used the first one correctly, but what would have happened if we had used the second by mistake?

$$72.0\ \text{in.} \times \frac{1\ \text{in.}}{2.54\ \text{cm}} = 28.3\ \text{in.}^2/\text{cm}$$

In this case, none of the units cancel. We get units of in.^{2}/cm because inches times inches is inches squared. Even though our calculator may be very good at arithmetic, we've got the wrong answer. *The factor-label method lets us know we have the wrong answer because the units are wrong!*

We will use the factor-label method extensively throughout this book to aid us in setting up the proper arithmetic in problems. In fact, we will see that it also helps us assemble the information we need to solve the problem. The following examples illustrate the method.

EXAMPLE 1.3
Applying the Factor-Label Method

Convert 3.25 m to millimeters (mm).

ANALYSIS: To clearly identify the problem, let's write the given quantity (with its units) on the left and the *units* of the desired answer on the right.

$$3.25\ \text{m} = ?\ \text{mm}$$

To solve this problem our tool will be a conversion factor that relates the unit meter to the unit millimeter. From the table of decimal multipliers, the prefix "milli" means "$\times\ 10^{-3}$," so we can write

$$1\ \text{mm} = 10^{-3}\ \text{m}$$

Notice that this relationship connects the units given to the units desired.

We now have all the information we need to solve the problem.

SOLUTION: From the relationship above, we can form two conversion factors.

$$\frac{1 \text{ mm}}{10^{-3} \text{ m}} \quad \text{and} \quad \frac{10^{-3} \text{ m}}{1 \text{ mm}}$$

We know we have to cancel the unit meter, so we need to multiply by a conversion factor with this unit in the denominator. Therefore, we select the one on the left as our tool. This gives

$$3.25 \cancel{\text{m}} \times \frac{1 \text{ mm}}{10^{-3} \cancel{\text{m}}} = 3.25 \times 10^3 \text{ mm}$$

Notice we have expressed the answer to three significant figures because that is how many there are in the given quantity, 3.25 m.

IS THE ANSWER REASONABLE? We know that millimeters are much smaller than meters, so 3.25 m must represent a lot of millimeters. Our answer, therefore, makes sense.

EXAMPLE 1.4
Applying the Factor-Label Method

A liter, which is slightly larger than a quart, is defined as 1 cubic decimeter (1 dm^3). How many liters are there in 1 cubic meter (1 m^3)?

ANALYSIS: Let's begin once again by stating the problem in equation form.

$$1 \text{ m}^3 = ? \text{ L}$$

Next, we assemble the tools. What relationships do we know that relate these various units? We are given the relationship between liters and cubic decimeters,

$$1 \text{ L} = 1 \text{ dm}^3 \tag{1.4}$$

From the table of decimal multipliers, we also know the relationship between decimeters and meters,

$$1 \text{ dm} = 0.1 \text{ m}$$

but we need a relationship between cubic units. Since units undergo the same kinds of operations numbers do, we simply cube each side of this equation (being careful to cube *both* the numbers and the units).

$$(1 \text{ dm})^3 = (0.1 \text{ m})^3$$

$$1 \text{ dm}^3 = 0.001 \text{ m}^3 \tag{1.5}$$

Notice how Equations 1.4 and 1.5 provide a path from the given units to those we seek. Such a path is always a necessary condition when we apply the factor-label method.

$$\text{m}^3 \xrightarrow[\text{Equation 1.5}]{} \text{dm}^3 \xrightarrow{\text{Equation 1.4}} \text{L}$$

Now we are ready to solve the problem.

SOLUTION: The first step is to eliminate the units m^3. We use Equation 1.5.

$$1 \cancel{\text{m}^3} \times \frac{1 \text{ dm}^3}{0.001 \cancel{\text{m}^3}} = 1000 \text{ dm}^3$$

Then we use Equation 1.4 to take us from dm^3 to L.

$$1000 \cancel{\text{dm}^3} \times \frac{1 \text{ L}}{1 \cancel{\text{dm}^3}} = 1000 \text{ L}$$

Thus, 1 m^3 = 1000 L.

Usually, when a problem involves the use of two or more conversion factors, they can be "strung together" in a "chain calculation" to avoid having to compute intermediate results. For example, this problem can be set up as follows.

$$1\ \cancel{m^3} \times \frac{1\ \cancel{dm^3}}{0.001\ \cancel{m^3}} \times \frac{1\ L}{1\ \cancel{dm^3}} = 1000\ L$$

IS THE ANSWER REASONABLE? One liter is about a quart. A cubic meter is about a cubic yard. Therefore, we expect a large number of liters in a cubic meter, so our answer seems reasonable. (Notice here that in our analysis we have approximated the quantities in the calculation in units of quarts and cubic yards, which may be more familiar than liters and cubic meters if you've been raised in the United States. We get a feel for the approximate magnitude of the answer using our familiar units and then relate this to the actual units of the problem.)

EXAMPLE 1.5
Applying the Factor-Label Method

Some mountain climbers are susceptible to high altitude pulmonary edema (HAPE), a life-threatening condition that causes fluid retention in the lungs. It can develop when a person climbs rapidly to heights greater than 2,500 meters (2.5×10^3 m). What is this distance expressed in feet?

ANALYSIS: The problem can be stated as

$$2500\ m = ?\ ft$$

We are converting a metric unit of length (the meter) into an English unit of length (the foot). The critical link between the two will be a metric-to-English length conversion. One of several sets of tools we can use is

$$1\ cm = 10^{-2}\ m \qquad \text{(from Table 1.4)}$$
$$1\ in. = 2.54\ cm \qquad \text{(from Table 1.6)}$$
$$1\ ft = 12\ in.$$

Notice how they provide a path from meters to centimeters to inches to feet.

SOLUTION: Now we apply the factor-label method by eliminating unwanted units to bring us to the units of the answer.

$$2.5 \times 10^3\ \cancel{m} \times \frac{1\ \cancel{cm}}{10^{-2}\ \cancel{m}} \times \frac{1\ \cancel{in.}}{2.54\ \cancel{cm}} \times \frac{1\ ft}{12\ \cancel{in.}} = 8.2 \times 10^3\ ft$$

Notice that if we were to stop after the first conversion factor, the units of the answer would be centimeters; if we stop after the second, the units would be inches, and after the third we get feet—the units we want. This time the answer has been rounded to two significant figures because that's how many there were in the measured distance. Notice that the numbers 12 and 2.54 do not affect the number of significant figures in the answer because they are exact numbers derived from definitions.

This is not the only way we could have solved this problem. Other sets of conversion factors could have been chosen. For example, we could have used 1 yd = 0.9144 m and 3 ft = 1 yd. Then the problem would have been set up as

$$2500\ \cancel{m} \times \frac{1\ \cancel{yd}}{0.9144\ \cancel{m}} \times \frac{3\ ft}{1\ \cancel{yd}} = 8200\ ft \qquad \text{(rounded correctly)}$$

Many problems that you meet, just like this one, have more than one path to the answer. There isn't necessarily any *one* correct way to set up the solution. *The important thing is for you to be able to reason your way through a problem and find some set of relationships that can take you from the given information to the answer.* The factor-label method can help you search for these relationships if you keep in mind the units that must be eliminated by cancellation.

IS THE ANSWER REASONABLE? Let's do some approximate arithmetic to get a feel for the size of the answer. A meter is slightly longer than a yard, so let's approximate the given distance, 2500 m, as 2500 yd. In 2500 yd, there are $3 \times 2500 = 7500$ ft. Since the meter is a bit longer than a yard, our answer should be a bit longer than 7500 ft, so the answer of 8200 ft seems to be reasonable.

Practice Exercise 7: Use the factor-label method to convert an area of 124 ft^2 to square meters. (Hint: What relationships would be required to convert feet to meters?)

Practice Exercise 8: Use the factor-label method to perform the following conversions: (a) 3.00 yd to inches, (b) 1.25 km to centimeters, (c) 3.27 mm to feet, and (d) 20.2 miles/gallon to kilometers/liter.

1.8 Density Is a Useful Intensive Property

In our earlier discussion of properties we noted that intensive properties are useful for identifying substances. One of the interesting things about extensive properties is that if you take the ratio of two of them, the resulting quantity is usually independent of sample size. In effect, the sample size cancels out and the calculated quantity becomes an intensive property. A useful property obtained this way is **density,** *which is defined as the ratio of an object's mass to its volume*. Using the symbols d for density, m for mass, and V for volume, we can express this mathematically as

$$d = \frac{m}{V} \qquad (1.6)$$

TOOLS
Density

Notice that to determine an object's density we make two measurements, mass and volume.

EXAMPLE 1.6
Calculating Density

A sample of blood completely fills an 8.20 cm^3 vial. The empty vial has a mass of 10.30 g. The vial has a mass of 18.91 g after being filled with blood. What is the density of blood?

ANALYSIS: This problem asks you to connect the mass and volume of blood with its density. The critical link between these quantities is the definition of density, given by Equation 1.6. (Without knowing this definition, you cannot solve the problem.) Equation 1.6 becomes the tool we'll use to obtain the answer.

SOLUTION: The volume of the blood equals the volume of the vial, 8.20 cm^3. The mass of the blood is the difference between the masses of the full and empty vials:

$$\text{Mass of blood} = 18.91\text{ g} - 10.30\text{ g} = 8.61\text{ g}$$

To determine the density we simply take the ratio of mass to volume.

$$\text{Density} = \frac{m}{V} = \frac{8.61\text{ g}}{8.20\text{ cm}^3} = 1.05\text{ g/cm}^3$$

This could also be written as

$$\text{Density} = 1.05\text{ g/mL}$$

because 1 cm^3 = 1 mL.

IS THE ANSWER REASONABLE? First, the answer has the correct units, so that's encouraging. In the calculation we are dividing 8.61 by 8.20, a number that is slightly smaller. The answer should be slightly larger than one, which it is, so a density of 1.05 g/cm^3 seems reasonable.

There is more mass in 1 cm^3 of gold than in 1 cm^3 of iron.

Although the density of water varies slightly with temperature, it is useful to remember the value 1.00 g/cm^3. It can be used if the water is near room temperature and only three (or fewer) significant figures are required.

Each pure substance has its own characteristic density (Table 1.7). Gold, for instance, is much more dense than iron. Each cubic centimeter of gold has a mass of 19.3 g, so its density is 19.3 g/cm^3. By comparison, the density of water is 1.00 g/cm^3 and the density of air at room temperature is about 0.0012 g/cm^3.

Most substances, such as the mercury in the bulb of a thermometer, expand slightly when they are heated, so the amount of matter packed into each cubic centimeter is less. Therefore, density usually decreases slightly with increasing temperature.[13] For solids and liquids the size of this change is small, as you can see from the data for water in Table 1.8. When only two or three significant figures are required, we can often ignore the variation of density with temperature.

TABLE 1.7

Densities of Some Common Substances in g/cm^3 at Room Temperature

Water	1.00
Aluminum	2.70
Iron	7.86
Silver	10.5
Gold	19.3
Glass	2.2
Air	0.0012

TABLE 1.8 Density of Water as a Function of Temperature

Temperature (°C)	Density (g/cm^3)
10	0.999700
15	0.999099
20	0.998203
25	0.997044
30	0.995646

Use density to relate a material's mass to its volume

A useful property of density is that it provides a way to convert between the mass and volume of a substance. It defines a relationship, which we will call an **equivalence**, between the amount of mass and its volume. For instance, the density of gold (19.3 g/cm^3) tells us that 19.3 g of the metal is equivalent to a volume of 1.00 cm^3. We express this relationship symbolically as

$$19.3 \text{ g gold} \Leftrightarrow 1.00 \text{ cm}^3 \text{ gold}$$

where we have used the symbol $\Leftrightarrow$ to mean "is equivalent to." (We can't really use an equals sign in this expression because grams can't *equal* cubic centimeters; one is a unit of mass and the other is a unit of volume.)

In setting up calculations by the factor-label method, an equivalence can be used to construct conversion factors just as equalities can. From the equivalence we have just written we can form two conversion factors:

$$\frac{19.3 \text{ g gold}}{1.00 \text{ cm}^3 \text{ gold}} \quad \text{and} \quad \frac{1.00 \text{ cm}^3 \text{ gold}}{19.3 \text{ g gold}}$$

The following example illustrates how we use density in calculations.

EXAMPLE 1.7
Calculations Using Density

Seawater has a density of about 1.03 g/mL. (a) What mass of seawater would fill a sampling vessel to a volume of 225 mL? (b) What is the volume, in milliliters, of 45.0 g of seawater?

ANALYSIS: For both parts of this problem, we are relating the mass of a material to its volume. Density is the critical link that we need between these two quantities. The given density tells us that *1.03 g of seawater is equivalent to 1.00 mL of seawater*, which we write as

$$1.03 \text{ g seawater} \Leftrightarrow 1.00 \text{ mL seawater}$$

[13] Liquid water behaves oddly. Its maximum density is at 4 °C, so when water at 0 °C is warmed, its density increases until the temperature reaches 4 °C. As the temperature is increased further the density of water gradually decreases.

From this relationship we can construct two conversion factors. These will be the tools we use to obtain the answers.

$$\frac{1.03 \text{ g seawater}}{1.00 \text{ mL seawater}} \quad \text{and} \quad \frac{1.00 \text{ mL seawater}}{1.03 \text{ g seawater}}$$

SOLUTION: (a) The question can be restated as 225 mL seawater ⇔ ? g seawater. We need to eliminate the unit *mL seawater*, so we choose the conversion factor on the left as our tool.

$$225 \cancel{\text{mL seawater}} \times \frac{1.03 \text{ g seawater}}{1.00 \cancel{\text{mL seawater}}} \Leftrightarrow 232 \text{ g seawater}$$

Thus, 225 mL of seawater has a mass of 232 g.

(b) The question is, 45.0 g seawater ⇔ ? mL seawater. This time we need to eliminate the unit *g seawater*, so we use the conversion factor on the right as our tool.

$$45.0 \cancel{\text{g seawater}} \times \frac{1.00 \text{ mL seawater}}{1.03 \cancel{\text{g seawater}}} \Leftrightarrow 43.7 \text{ mL seawater}$$

Thus, 45.0 g of seawater has a volume of 43.7 mL.

ARE THE ANSWERS REASONABLE? Notice that the density tells us that 1 mL of seawater has a mass of slightly more than 1 g. So for part (a), we might expect that 225 mL of seawater should have a mass slightly more than 225 g. Our answer, 232 g, is reasonable. For part (b), 45 g of seawater should have a volume not too far from 45 mL, so our answer of 43.7 mL is the right size.

Practice Exercise 9: A gold-colored metal object has a mass of 365 g and a volume of 22.12 cm^3. Is the object composed of pure gold? (Hint: How does the density of the object compare with that of pure gold?)

Practice Exercise 10: A certain metal alloy has a density of 12.6 g/cm^3. How many pounds would 0.822 ft^3 of this alloy weigh? (Hint: What is the density of the alloy in units of lb/ft^3?)

Practice Exercise 11: An ocean-dwelling dinosaur was estimated to have had a body volume of 1.38×10^6 cm^3. The animal's mass when alive was estimated at 1.24×10^6 g. What is its density?

Practice Exercise 12: The density of diamond is 3.52 g/cm^3. What is the volume in cubic centimeters of a 1 carat diamond, which has a mass of 200 mg? (Assume three significant figures.)

Conclusions must be drawn from reliable measurements

We saw earlier that substances can be identified by their properties. If we are to rely on properties such as density for identification of substances, it is very important that our measurements be reliable. We must have some idea of what the measurement's accuracy and precision are.

The importance of accuracy is obvious. If we have no confidence that our measured values are close to the true values, we certainly cannot trust any conclusions that are based on the data we have collected.

Precision of measurements can be equally important. For example, suppose we had a gold wedding ring and we wanted to determine whether or not the gold was 24 carat. We could determine the mass of the ring, and then its volume, and compute the density of the ring. We could then compare our experimental density with the density of 24 carat gold (which is 19.3 g/mL). Suppose the ring had a volume of 1.0 mL and the ring had a mass of 18 g, as measured using a graduated cup measure and a kitchen scale. The density of the ring would then be 18 g/mL, to the correct number of significant figures. Could we conclude that the ring was made of 24 carat gold? We know the density to only two significant figures. The experimental density could be as low as 17 g/mL or as high as 19 g/mL, which means the ring *could* be 24 carat gold—or it could be 22 carat gold (which has a density of around 17.7 to 17.8 g/mL) or maybe even 18 carat gold (which has a density up to 16.9 g/mL).

Suppose we now measure the mass of the ring with a laboratory balance capable of measurements to the nearest ±0.001 g and obtain a mass of 18.153 g. We measure the volume using volumetric glassware and find a volume of 1.03 mL. The density is 17.6 g/mL to the correct number of significant figures. The difference between this density and the density of 24 carat gold is 19.3 g/mL − 17.6 g/mL = 1.7 g/mL. This is considerably larger than the uncertainty in the experimental density (which is about ±0.1 g/mL). We can be reasonably confident that the ring is not 24 carat gold, and in fact the measurements point toward the ring being composed of 22 carat gold.

To trust conclusions drawn from measurements, we must be sure the measurements are accurate and that they are of sufficient precision to be meaningful. This is a key consideration in designing experiments.

SUMMARY

Chemistry and the Scientific Method. **Chemistry** is a science that studies the properties and composition of **matter**, which is anything that has **mass** and occupies space. It employs the **scientific method** in which **observations** are used to collect **empirical facts,** or **data,** that can be summarized in **scientific laws. Models** of nature begin as **hypotheses** that mature into **theories** when they survive repeated testing. According to the **atomic theory,** matter is composed of **atoms** that combine to form more complex substances, many of which consist of **molecules** composed of two or more atoms.

Elements, Compounds, and Mixtures. An **element,** which is identified by its **chemical symbol,** cannot be decomposed into something simpler by a **chemical reaction.** Elements combine in fixed proportions to form **compounds.** Elements and compounds are **pure substances** that may be combined in *varying* proportions to give **mixtures.** If a mixture has two or more **phases,** it is **heterogeneous.** A one-phase **homogeneous** mixture is called a **solution.** Formation or separation of a mixture into its components can be accomplished by a **physical change,** which doesn't alter the chemical composition of the substances involved. Formation or decomposition of a compound takes place by a **chemical change** that changes the chemical makeup of the substances involved.

Properties of Materials. **Physical properties** are measured without changing the chemical composition of a sample. **Solid, liquid,** and **gas** are the most common **states of matter.** Their properties can be related to the different ways the individual atomic-size particles are organized. A **chemical property** describes a chemical reaction a substance undergoes. **Intensive properties** are independent of sample size; **extensive properties** depend on sample size.

Units of Measurement. **Qualitative observations** lack numerical information, whereas **quantitative observations** require numerical measurements. The units used for scientific measurements are based on the set of seven **SI base units** which can be combined to give various **derived units.** These all can be scaled to larger or smaller sized units by applying **decimal multiplying factors.** In the laboratory we routinely measure length, volume, mass, and temperature. Convenient units for length and volume are, respectively, **centimeters** or **millimeters,** and **liters** or **milliliters. Mass** is a measure of the amount of matter in an object and differs from weight. Mass is measured with a **balance** and is expressed in units of **kilograms** or **grams.** Temperature is measured in units of **degrees Celsius** (or **Fahrenheit**) using a thermometer. For many calculations, temperature must be expressed in **kelvins (K).** The zero point on the **Kelvin temperature scale** is called **absolute zero.**

Significant Figures. The **precision** of a measured quantity is revealed by the number of **significant figures** that it contains, which equals the number of digits known for sure plus the first one that possesses some uncertainty. Measured values are **precise** if they contain many significant figures and therefore differ from each other by small amounts. A measurement is **accurate** if its value lies very close to the true value. When measurements are combined in calculations, rules help us determine the correct number of significant figures in the answer (see below). **Exact numbers** are considered to have an infinite number of significant figures.

Factor-Label Method. The **factor-label method** is based on the ability of units to undergo the same mathematical operations as numbers. **Conversion factors** are constructed from *valid relationships* between units. These relationships can be either equalities or **equivalencies** (indicated by the symbol ⇔) between units. Unit cancellation serves as a guide to the use of conversion factors and aids us in correctly setting up the arithmetic for a problem.

Density. **Density** is an intensive property equal to the ratio of a sample's mass to its volume. Besides serving as a means for identifying substances, density provides a conversion factor that relates mass to volume.

TOOLS FOR PROBLEM SOLVING

In this chapter you learned to apply the following concepts as tools in solving problems. Study each one carefully so that you know what each is used for. When faced with solving a problem, recall what each tool does and consider whether it will be helpful in finding a solution. This will aid you in selecting the tools you need.

SI Prefixes *(Table 1.4, page 12)* We use the prefixes to create larger and smaller units. They are also used to create conversion factors for converting between differently sized units. Be sure you are familiar with the ones in bold colored type in Table 1.4.

Units in laboratory measurements *(pages 13-17)* Often we must convert among units commonly used for laboratory measurements.

Length: $1 \text{ m} = 100 \text{ cm} = 1000 \text{ mm}$

Volume: $1 \text{ L} = 1000 \text{ mL} = 1000 \text{ cm}^3$

Temperature conversions *(pages 16-17)* Use Equations 1.1 and 1.2 to convert between temperature scales.

$$t_F = \left(\frac{9\ °\text{F}}{5\ °\text{C}}\right)t_C + 32\ °\text{F} \qquad T_K = (t_C + 273.15\ °\text{C})\left(\frac{1\ \text{K}}{1\ °\text{C}}\right)$$

Add 273.15 to Celsius temperature to obtain the Kelvin temperature.

Rules for counting significant figures in a number *(page 19)* To gauge the quality of a measurement, we must know the number of significant figures it contains:

- All nonzero digits are significant.
- Zeros to the right of the decimal are significant if they follow a nonzero digit.
- Zeros between significant digits are significant.
- Zeros that are to the *left* of the first nonzero digit are not significant.
- Zeros on the end of a number without a decimal point are assumed not to be significant. (To avoid confusion, scientific notation should be used.)

Rules for arithmetic and significant figures *(pages 20-21)* We use these rules in almost every numerical problem to obtain the correct number of significant figures in the answer.

Multiplication and division: Round the answer to the same number of significant figures as the least precise factor.

Addition and subtraction: Round the answer to match the same number of decimal places as the quantity with the fewest number of decimal places.

Exact numbers: Exact numbers, such as those that arise from definitions, do not affect the number of significant figures in the result of a calculation.

Density *(page 25)* The density, d, relates mass, m, and volume, V, for a substance.

$$d = \frac{m}{V}$$

Density provides an equivalence between mass and volume, from which we can construct conversion factors to convert between mass and volume for a substance.

QUESTIONS, PROBLEMS, AND EXERCISES

Answers to problems whose numbers are printed in color are given in Appendix B. More challenging problems are marked with asterisks. ILW = Interactive Learningware solution is available at *www.wiley.com/college/brady*. OH = an Office Hours video is available for this problem.

REVIEW QUESTIONS

Introduction; The Scientific Method

1.1 After some thought, give two reasons why a course in chemistry will benefit *you* in the pursuit of your particular major.

1.2 What steps are involved in the scientific method?

1.3 What is the difference between (a) a law and a theory, (b) an observation and a conclusion, (c) an observation and data?

Properties of Substances

OH **1.4** Define *matter*. Which of the following are examples of matter? (a) air, (b) a pencil, (c) a cheese sandwich, (d) a squirrel, (e) your mother

1.5 What is *a physical property?* What is *a chemical property?* What is the chief distinction between physical and chemical properties? Define the terms *intensive property* and *extensive property.* Give two examples of each.

1.6 "A sample of calcium (an electrically conducting white metal that is shiny, relatively soft, melts at 850 °C, and boils at 1440 °C) was placed into liquid water that was at 25 °C. The calcium reacted slowly with the water to give bubbles of gaseous hydrogen and a solution of the substance calcium hydroxide." In this description, what physical properties and what chemical properties are described?

OH **1.7** In places like Saudi Arabia, freshwater is scarce and is recovered from seawater. When seawater is boiled, the water evaporates and the steam can be condensed to give pure water that people can drink. If all the water is evaporated, solid salt is left behind. Are the changes described here chemical or physical?

1.8 Name the three states of matter.

Elements, Compounds, and Mixtures

1.9 Define (a) element, (b) compound, (c) mixture, (d) homogeneous, (e) heterogeneous, (f) phase, and (g) solution.

1.10 What is the chemical symbol for each of the following elements? (a) chlorine, (b) sulfur, (c) iron, (d) silver, (e) sodium, (f) phosphorus, (g) iodine, (h) copper, (i) mercury, (j) calcium

1.11 What is the name of each of the following elements? (a) K, (b) Zn, (c) Si, (d) Sn, (e) Mn, (f) Mg, (g) Ni, (h) Al, (i) C, (j) N

SI Units

1.12 Why must measurements always be written with units?

1.13 What is the only SI base unit that includes a decimal prefix?

1.14 What is the meaning of each of the following prefixes? (a) centi-, (b) milli-, (c) kilo-, (d) micro-, (e) nano-, (f) pico-, (g) mega-

1.15 What abbreviation is used for each of the prefixes named in Question 1.14?

1.16 What reference points do we use in calibrating the scale of a thermometer? What temperature on the Celsius scale do we assign to each of these reference points?

OH **1.17** In each pair, which is larger: (a) A Fahrenheit degree or a Celsius degree? (b) A Celsius degree or a kelvin? (c) A Fahrenheit degree or a kelvin?

Significant Figures; the Factor-Label Method

1.18 Define the term *significant figures.*

1.19 What is the difference between *accuracy* and *precision*?

1.20 Suppose a length had been reported to be 31.24 cm. What is the minimum uncertainty implied in this measurement?

1.21 Suppose someone suggested using the fraction $\frac{3\text{ yd}}{1\text{ ft}}$ as a conversion factor to change a length expressed in feet to its equivalent in yards. What is wrong with this conversion factor? Can we construct a valid conversion factor relating centimeters to meters from the equation 1 cm = 1000 m? Explain your answer.

1.22 In 1 hour there are 3600 seconds. By what conversion factor would you multiply 250 seconds to convert it to hours? By what conversion factor would you multiply 3.84 hours to convert it to seconds?

1.23 If you were to convert the measured length 4.165 ft to yards by multiplying by the conversion factor (1 yd/3 ft), how many significant figures should the answer contain? Why?

Density

1.24 Write the equation that defines density. Identify the symbols in the equation.

1.25 Silver has a density of 10.5 g cm^{-3}. Express this as an equivalence between mass and volume for silver. Write two conversion factors that can be formed from this equivalence for use in calculations.

REVIEW PROBLEMS

SI Prefixes

1.26 What number should replace the question mark in each of the following?
(a) 1 cm = ? m
(b) 1 km = ? m
(c) 1 m = ? pm
(d) 1 dm = ? m
(e) 1 g = ? kg
(f) 1 cg = ? g

1.27 What numbers should replace the question marks below?
(a) 1 nm = ? m
(b) 1 μg = ? g
(c) 1 kg = ? g
(d) 1 Mg = ? g
(e) 1 mg = ? g
(f) 1 dg = ? g

Temperature Conversions

1.28 Perform the following conversions.
(a) 50 °C to °F
(b) 10 °C to °F
(c) 25.5 °F to °C
(d) 49 °F to °C
(e) 60 °C to K
(f) −30 °C to K

1.29 Perform the following conversions.
(a) 96 °F to °C
(b) −6 °F to °C
(c) −55 °C to °F
(d) 273 K to °C
(e) 299 K to °C
(f) 40 °C to K

1.30 A healthy dog has a temperature ranging from 37.2 to 39.2 °C. Is a dog with a temperature of 103.5 °F within normal range?

1.31 The coldest permanently inhabited place on earth is the Siberian village of Oymyakon in Russia. In 1964 the temperature reached a shivering −96 °F! What is this temperature in °C?

1.32 Estimates of the temperature at the core of the sun range from 10 megakelvins to 25 megakelvins. What is this range in °C and °F?

1.33 Natural gas is mostly methane, a substance that boils at a temperature of 111 K. What is its boiling point in °C and °F?

1.34 Helium has the lowest boiling point of any liquid. It boils at 4 K. What is its boiling point in °C?

1.35 The atomic bomb detonated over Hiroshima, Japan, at the end of World War II raised the temperature on the ground below to about 6000 K. Is this hot enough to melt concrete? (Concrete melts at 2000 °C.)

Significant Figures

1.36 How many significant figures do the following measured quantities have?
(a) 37.53 cm
(b) 37.240 cm
(c) 202.0 g
(d) 0.00024 kg
(e) 0.07080 m
(f) 2400 mL

1.37 How many significant figures do the following measured quantities have?
(a) 0.0230 g
(b) 105.303 m
(c) 0.007 kg
(d) 614.00 mg
(e) 10 L
(f) 3.8105 mm

OH **1.38** Perform the following arithmetic and round off the answers to the correct number of significant figures. Include the correct units with the answers.
(a) 0.0023 m × 315 m
(b) 84.25 kg – 0.01075 kg
(c) (184.45 g − 94.45 g)/(31.4 mL − 9.9 mL)
(d) (23.4 g + 102.4 g + 0.003 g)/(6.478 mL)
(e) (313.44 cm – 209.1 cm) × 8.2234 cm

1.39 Perform the following arithmetic and round off the answers to the correct number of significant figures. Include the correct units with the answers.
(a) 3.58 g/1.739 mL
(b) 4.02 mL + 0.001 mL
(c) (22.4 g − 8.3 g)/(1.142 mL − 0.002 mL)
(d) (1.345 g + 0.022 g)/(13.36 mL – 8.4115 mL)
(e) (74.335 m – 74.332 m)/(4.75 s × 1.114 s)

Unit Conversions by the Factor-Label Method

OH **1.40** Perform the following conversions.
(a) 32.0 dm/s to km/hr (d) 137.5 mL to L
(b) 8.2 mg/mL to μg/L (e) 0.025 L to mL
(c) 75.3 mg to kg (f) 342 pm^2 to dm^2

1.41 Perform the following conversions.
(a) 92 dL to μm^3 (d) 230 km^3 to m^3
(b) 22 ng to μg (e) 87.3 cm s^{-2} to km hr^{-2}
(c) 83 pL to nL (f) 238 mm^2 to nm^2

1.42 Perform the following conversions. If necessary, refer to Tables 1.4 and 1.6.
(a) 36 in. to cm (d) 1 cup (8 oz) to mL
(b) 5.0 lb to kg (e) 55 mi/hr to km/hr
(c) 3.0 qt to mL (f) 50.0 mi to km

1.43 Perform the following conversions. If necessary, refer to Tables 1.4 and 1.6.
(a) 250 mL to qt (d) 1.75 L to fluid oz
(b) 3.0 ft to m (e) 35 km/hr to mi/hr
(c) 1.62 kg to lb (f) 80.0 km to mi

1.44 Perform the following conversions.
(a) 8.4 ft^2 to cm^2 (b) 223 mi^2 to km^2 (c) 231 ft^3 to cm^3

1.45 Perform the following conversions.
(a) 2.4 yd^2 to m^2 (b) 8.3 in.^{2} to mm^2 (c) 9.1 ft^3 to L

1.46 The human stomach can expand to hold up to 4.2 quarts of food. A pistachio nut has a volume of about 0.9 mL. Use this information to estimate the maximum number of pistachios that can be eaten in one sitting.

1.47 In the movie *Cool Hand Luke* (1967), Luke wagers that he can eat 50 eggs in one hour. The prisoners and guards bet against him, saying, "Fifty eggs gotta weigh a good six pounds. A man's gut can't hold that." A chewed, peeled chicken egg has a volume of approximately 53 mL. If Luke's stomach has a volume of 4.2 quarts, does he have any chance of winning the bet?

ILW **1.48** The winds in a hurricane can reach almost 200 miles per hour. What is this speed in meters per second? (Assume three significant figures.)

1.49 A bullet is fired at a speed of 2435 ft/s. What is this speed expressed in kilometers per hour?

1.50 A bullet leaving the muzzle of a pistol was traveling at a speed of 2230 feet per second. What is this speed in miles per hour?

1.51 On average, water flows over Niagara Falls at a rate of 2.05 × 10^5 cubic feet per second. One cubic foot of water weighs 62.4 lb. Calculate the rate of water flow in tons of water per day. (1 ton = 2000 lb.)

1.52 The brightest star in the night sky in the northern hemisphere is Sirius. Its distance from earth is estimated to be 8.7 light-years. A light-year is the distance light travels in one year. Light travels at a speed of 3.00 × 10^8 m/s. Calculate the distance from earth to Sirius in miles. (1 mi = 5280 ft.)

1.53 One degree of latitude on the earth's surface equals 60.0 nautical miles. One nautical mile equals 1.151 statute miles. (A *statute mile* is the distance over land that we normally associate with the unit mile.) Calculate the circumference of the earth in statute miles.

1.54 The deepest point in the earth's oceans is found in the Mariana Trench, a deep crevasse located about 1000 miles southeast of Japan beneath the Pacific Ocean. Its maximum depth is 6033.5 fathoms. One fathom is defined as 6 feet. Calculate the depth of the Mariana Trench in meters.

1.55 At sea level, our atmosphere exerts a pressure of about 14.7 lb/in.^{2}, which means that each square inch of your body experiences a force of 14.7 lb from the air that surrounds you. As you descend below the surface of the ocean, the pressure produced by the seawater increases by about 14.7 lb/in.^{2} for every 10 meters of depth. In the preceding problem you calculated the maximum depth of the Mariana Trench, located in the Pacific Ocean. What is the approximate pressure in pounds per square inch and in tons per square inch exerted by the sea at the deepest point of the trench? (1 ton = 2000 lb.)

Density

1.56 A sample of kerosene weighs 36.4 g. Its volume was measured to be 45.6 mL. What is the density of the kerosene?

1.57 A block of magnesium has a mass of 14.3 g and a volume of 8.46 cm^3. What is the density of magnesium in g/cm^3?

OH **1.58** Acetone, the solvent in some nail polish removers, has a density of 0.791 g/mL. What is the volume of 25.0 g of acetone?

1.59 A glass apparatus contains 26.223 g of water when filled at 25 °C. At this temperature, water has a density of 0.99704 g/mL. What is the volume of this apparatus?

1.60 Chloroform, a chemical once used as an anesthetic, has a density of 1.492 g/mL. What is the mass in grams of 185 mL of chloroform?

1.61 Gasoline has a density of about 0.65 g/mL. How much does 34 L (approximately 18 gallons) weigh in kilograms? In pounds?

ILW **1.62** A graduated cylinder was filled with water to the 15.0 mL mark and weighed on a balance. Its mass was 27.35 g. An object made of silver was placed in the cylinder and completely submerged in the water. The water level rose to 18.3 mL. When reweighed, the cylinder, water, and silver object had a total mass of 62.00 g. Calculate the density of silver.

1.63 Titanium is a metal used to make golf clubs. A rectangular bar of this metal measuring 1.84 cm × 2.24 cm × 2.44 cm was found to have a mass of 45.7 g. What is the density of titanium?

1.64 The space shuttle uses liquid hydrogen as its fuel. The external fuel tank used during takeoff carries 227,641 lb of hydrogen with a volume of 385,265 gallons. Calculate the density of liquid hydrogen in units of g/mL. (Express your answer to three significant figures.)

1.65 Some time ago, a U.S. citizen traveling in Canada observed that the price of regular gasoline was 0.959 Canadian dollars per liter. The exchange rate at the time was 1.142 Canadian dollars per one U.S. dollar. Calculate the price of the Canadian gasoline in units of U.S. dollars per gallon. (Just the week before, the traveler had paid $2.249 per gallon in the United States.)

ADDITIONAL EXERCISES

OH **1.66** You are the science reporter for a daily newspaper and your editor asks you to write a story based on a report in the scientific literature. The report states that analysis of the sediments in Hausberg Tarn (elevation 4350 m) on the side of Mount Kenya (elevation 4600–4700 m) shows that the average temperature of the water rose by 4.0 °C between 350 BC and AD 450. Your editor tells you that she wants all the data expressed in the English system of units. Make the appropriate conversions.

1.67 An astronomy website states that neutron stars have a density of 1.00×10^8 tons per cubic centimeter. The site does not specify whether "tons" means metric tons (1 metric ton = 1000 kg) or English tons (1 English ton = 2000 pounds). How many grams would one teaspoon of a neutron star weigh, if the density were in metric tons per cm^3? How many grams would the teaspoon weigh if the density were in English tons per cm^3? (One teaspoon is approximately 4.93 mL.)

1.68 The star Arcturus is 3.50×10^{14} km from the earth. How many days does it take for light to travel from Arcturus to earth? What is the distance to Arcturus in light-years? One light-year is the distance light travels in one year (365 days); light travels at a speed of 3.00×10^8 m/s.

1.69 A pycnometer is a glass apparatus used for accurately determining the density of a liquid. When dry and empty, a certain pycnometer had a mass of 27.314 g. When filled with distilled water at 25.0 °C, it weighed 36.842 g. When filled with chloroform (a liquid once used as an anesthetic before its toxic properties were known), the apparatus weighed 41.428 g. At 25.0 °C, the density of water is 0.99704 g/mL. (a) What is the volume of the pycnometer? (b) What is the density of chloroform?

1.70 Radio waves travel at the speed of light, 3.00×10^8 m/s. If you were to broadcast a question to an astronaut on the moon, which is 239,000 miles from earth, what is the minimum time that you would have to wait to receive a reply?

1.71 Suppose you have a job in which you earn $4.50 for each 30 minutes that you work.

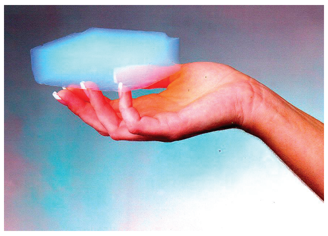
(a) Express this information in the form of an equivalence between dollars earned and minutes worked.

(b) Use the equivalence defined in (a) to calculate the number of dollars earned in 1 hr 45 min.

(c) Use the equivalence defined in (a) to calculate the number of minutes you would have to work to earn $17.35.

1.72 When an object floats in water, it displaces a volume of water that has a weight equal to the weight of the object. If a ship has a weight of 4255 tons, how many cubic feet of seawater will it displace? Seawater has a density of 1.025 g cm^{-3}; 1 ton = 2000 lb.

1.73 Aerogel or "solid smoke" is a novel material that is made of silicon dioxide, like glass, but is a thousand times less dense than glass because it is extremely porous. Material scientists at NASA's Jet Propulsion Laboratory created the lightest aerogel ever in 2002, with a density of 0.00011 pounds per cubic inch. The material was used for thermal insulation in the 2003 Mars Exploration Rover. If the maximum space for insulation in the spacecraft's hull was 2510 cm^3, what mass (in grams) did the aerogel insulation add to the spacecraft?



Aerogel. *(NASA/JPL.)*

1.74 A liquid known to be either ethanol (ethyl alcohol) or methanol (methyl alcohol) was found to have a density of 0.798 ± 0.001 g/mL. Consult the *Handbook of Chemistry and Physics* to determine which liquid it is. What other measurements could help to confirm the identity of the liquid?

1.75 An unknown liquid was found to have a density of 69.22 lb/ft^3. The density of ethylene glycol (the liquid used in antifreeze) is 1.1088 g/mL. Could the unknown liquid be ethylene glycol?

1.76 When an object is heated to a high temperature, it glows and gives off light. The color balance of this light depends on the temperature of the glowing object. Photographic lighting is described, in terms of its color balance, as a temperature in kelvins. For example, a certain electronic flash gives a color balance (called color temperature) rated at 5800 K. What is this temperature expressed in °C?

OH *__1.77__ There exists a single temperature at which the value reported in °F is numerically the same as the value reported in °C. What is that temperature?

*__1.78__ In the text, the Kelvin scale of temperature is defined as an absolute scale in which one Kelvin degree unit is the same size as one Celsius degree unit. A second absolute temperature scale exists called the Rankine scale. On this scale, one Rankine degree unit (°R) is the same size as one Fahrenheit degree unit. (a) What is the only temperature at which the Kelvin and Rankine scales possess the same numerical value? Explain your answer. (b) What is the boiling point of water expressed in °R?

*__1.79__ Density measurements can be used to analyze mixtures. For example, the density of solid sand (without air spaces) is about 2.84 g/mL. The density of gold is 19.3 g/mL. If a 1.00 kg sample of sand containing some gold has a density of 3.10 g/mL (without air spaces), what is the percentage of gold in the sample?

*__1.80__ An artist's statue has a surface area of 14.6 ft^2. The artist plans to apply gold plate to the statue and wants the coating to be 2.50 μm thick. If the price of gold were $625.10 per troy ounce, how much would it cost to give the statue its gold coating? (1 troy ounce = 31.1035 g; the density of gold is 19.3 g/mL.)

*__1.81__ A cylindrical metal bar has a diameter of 0.753 cm and a length of 2.33 cm. It has a mass of 8.423 g. Calculate the density of the metal in the units lb/ft^3.

1.82 What is the volume in cubic millimeters of a 3.54 carat diamond, given that the density of the diamond is 3.51 g/mL? (1 carat = 200 mg.)

1.83 Because of the serious consequences of lead poisoning, the Federal Centers for Disease Control in Atlanta has set a threshold of concern for lead levels in children's blood. This threshold was based on a study that suggested that lead levels in blood as low as *10 micrograms of lead per deciliter of blood* can result in subtle effects of lead toxicity. Suppose a child had a lead level in her blood of 2.5×10^{-4} grams of lead per liter of blood. Is this person in danger of exhibiting the effects of lead poisoning?

*__1.84__ Gold has a density of 19.31 g cm^{-3}. How many grams of gold are required to provide a gold coating 0.500 mm thick on a ball bearing having a diameter of 2.000 mm?

*__1.85__ A Boeing 747 jet airliner carrying 568 people burns about 5.0 gallons of jet fuel per mile. What is the rate of fuel consumption in units of miles per gallon per person? Is this better or worse than the rate of fuel consumption in an automobile carrying two people that gets 21.5 miles per gallon? If the airliner were making the 3470 mile trip from New York to London, how many pounds of jet fuel would be consumed? (Jet fuel has a density of 0.803 g/mL.)

EXERCISES IN CRITICAL THINKING

1.86 A homogeneous solution is defined as a uniform mixture consisting of a single phase. With our vastly improved abilities to "see" smaller and smaller particles, down to the atomic level, present an argument for the proposition that all mixtures are heterogeneous. Present the argument that the ability to observe objects as small as an atom has no effect on the definitions of heterogeneous and homogeneous.

1.87 Find two or more websites that give the values for each of the seven base SI units. Keeping in mind that not all websites provide reliable information, which website do you believe provides the most reliable values? Justify your answer.

1.88 Reference books such as the *Handbook of Chemistry and Physics* report the specific gravities of substances instead of their densities. Find the definition of specific gravity and discuss the relative merits of specific gravity and density in terms of their usefulness as a physical property.

1.89 A student used a graduated cylinder having volume markings every 2 mL to carefully measure 100 mL of water for an experiment. A fellow student said that by reporting the volume as "100 mL" in her lab notebook, she was only entitled to one significant figure. She disagreed. Why did her fellow student say the reported volume had only one significant figure? Considering the circumstances, how many significant figures are in her measured volume? Justify your answer.

1.90 Download a table of data for the density of water between its freezing and boiling points. Use a spreadsheet program to plot (a) the density of water versus temperature and (b) the volume of a kilogram of water versus temperature. Interpret the significance of these plots.

1.91 List the physical and chemical properties mentioned in this chapter. What additional physical and chemical properties can you think of to extend the list?