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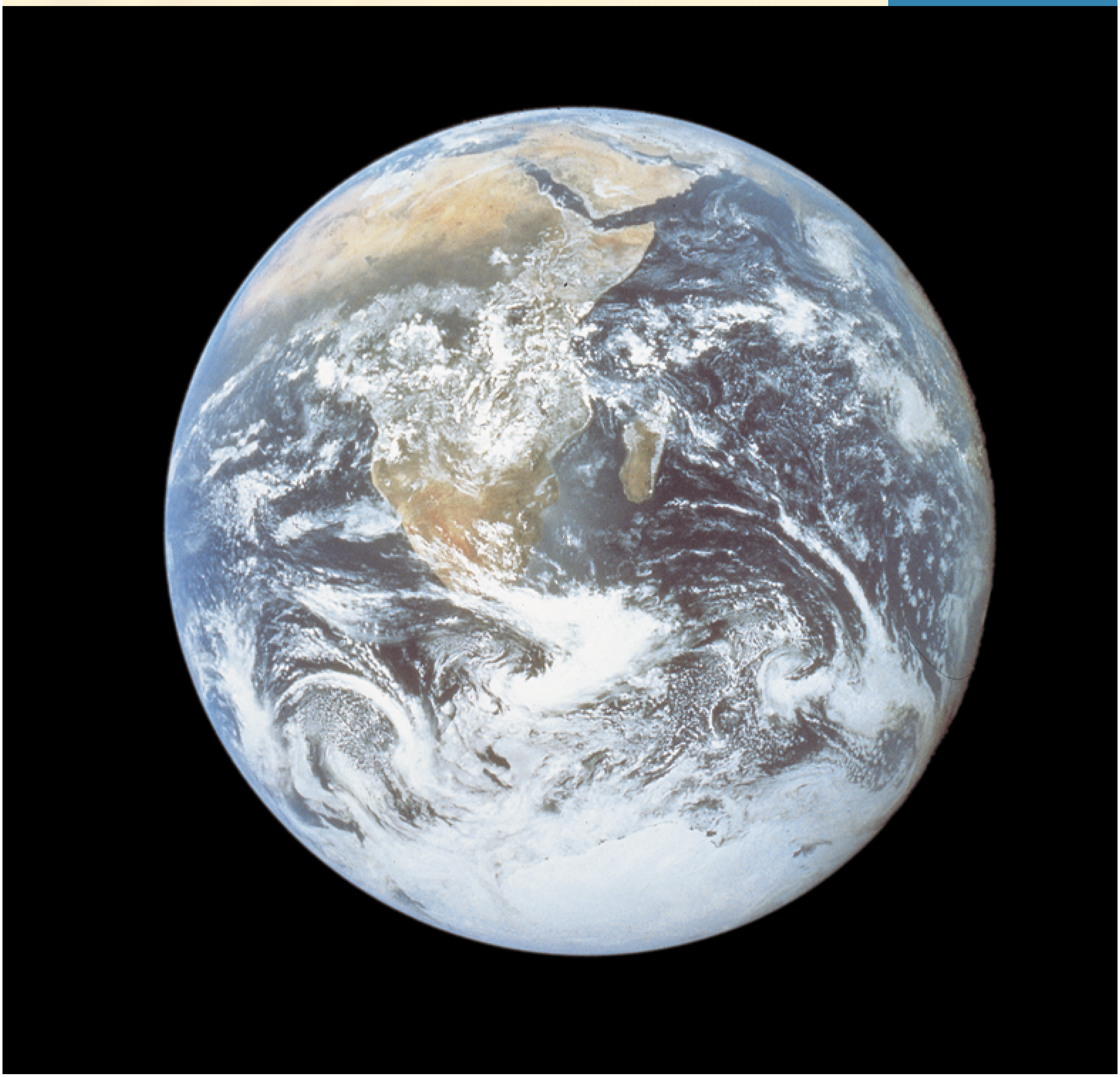
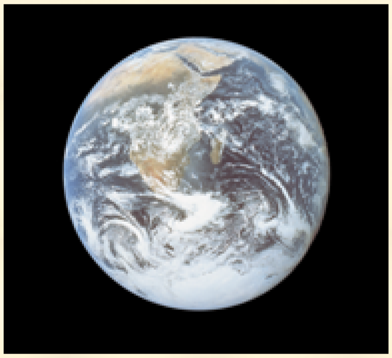
The Foundations of Chemistry





The earth is a huge chemical system, including innumerable reactions taking place constantly, with some energy input from sunlight. The earth serves as the source of raw materials for all human activities as well as the depository for the products of these activities. Maintaining life on the planet requires understanding and intelligent use of these resources. Scientists can provide important information about the processes, but each of us must share in the responsibility for our environment.

OUTLINE

OBJECTIVES

After you have studied this chapter, you should be able to

- *Use the basic vocabulary of matter and energy*
- *Distinguish between chemical and physical properties and between chemical and physical changes*
- *Recognize various forms of matter: homogeneous and heterogeneous mixtures, substances, compounds, and elements*
- *Apply the concept of significant figures*
- *Apply appropriate units to describe the results of measurement*
- *Use the unit factor method to carry out conversions among units*
- *Describe temperature measurements on various common scales, and convert between these scales*
- *Carry out calculations relating temperature change to heat absorbed or liberated*

Thousands of practical questions are studied by chemists. A few of them are

How can we modify a useful drug so as to improve its effectiveness while minimizing any harmful or unpleasant side effects?

How can we develop better materials to be used as synthetic bone for replacement surgery?

Which substances could help to avoid rejection of foreign tissue in organ transplants?

What improvements in fertilizers or pesticides can increase agricultural yields? How can this be done with minimal environmental danger?

How can we get the maximum work from a fuel while producing the least harmful emissions possible?

Which really poses the greater environmental threat—the burning of fossil fuels and its contribution to the greenhouse effect and climatic change, or the use of nuclear power and the related radiation and disposal problems?

How can we develop suitable materials for the semiconductor and microelectronics industry? Can we develop a battery that is cheaper, lighter, and more powerful?

What changes in structural materials could help to make aircraft lighter and more economical, yet at the same time stronger and safer?

What relationship is there between the substances we eat, drink, or breathe and the possibility of developing cancer? How can we develop substances that are effective in killing cancer cells preferentially over normal cells?

Can we economically produce fresh water from sea water for irrigation or consumption?

How can we slow down unfavorable reactions, such as the corrosion of metals, while speeding up favorable ones, such as the growth of foodstuffs?

Chemistry touches almost every aspect of our lives, our culture, and our environment. Its scope encompasses the air we breathe, the food we eat, the fluids we drink, our clothing, dwellings, transportation and fuel supplies, and our fellow creatures.

Enormous numbers of chemical reactions are necessary to produce a human embryo (here at 10 weeks, 6 cm long).

> Chemistry is the science that describes matter—its properties, the changes it undergoes, and the energy changes that accompany those processes.

Matter includes everything that is tangible, from our bodies and the stuff of our everyday lives to the grandest objects in the universe. Some call chemistry the central science. It rests on the foundation of mathematics and physics and in turn underlies the life sciences—biology and medicine. To understand living systems fully, we must first understand the chemical reactions and chemical influences that operate within them. The chemicals of our bodies profoundly affect even the personal world of our thoughts and emotions.

No one can be expert in all aspects of such a broad science as chemistry. Sometimes we arbitrarily divide the study of chemistry into various branches. Carbon is very versatile in its bonding and behavior and is a key element in many substances that are essential to life. All living matter contains carbon combined with hydrogen. The chemistry of compounds of carbon and hydrogen is called **organic chemistry.** (In the early days of chemistry, living matter and inanimate matter were believed to be entirely different. We now know that many of the compounds found in living matter can be made from nonliving, or "inorganic," sources. Thus, the terms "organic" and "inorganic" have different meanings than they did originally.) The study of substances that do not contain carbon combined with hydrogen is called **inorganic chemistry.** The branch of chemistry that is concerned with the detection or identification of substances present in a sample (*qualitative analysis*) or with the amount of each that is present (*quantitative analysis*) is called **analytical chemistry. Physical chemistry** applies the mathematical theories and methods of physics to the properties of matter and to the study of chemical processes and the accompanying energy changes. As its name suggests, **biochemistry** is the study of the chemistry of processes in living organisms. Such divisions are arbitrary, and most chemical studies involve more than one of these traditional areas of chemistry. The principles you will learn in a general chemistry course are the foundation of all branches of chemistry.

We understand simple chemical systems well; they lie near chemistry's fuzzy boundary with physics. They can often be described exactly by mathematical equations. We fare less well with more complicated systems. Even where our understanding is fairly thorough, we must make approximations, and often our knowledge is far from complete. Each year researchers provide new insights into the nature of matter and its interactions. As chemists find answers to old questions, they learn to ask new ones. Our scientific knowledge has been described as an expanding sphere that, as it grows, encounters an ever-enlarging frontier.

In our search for understanding, we eventually must ask fundamental questions, such as the following:

How do substances combine to form other substances? How much energy is involved in changes that we observe?

How is matter constructed in its intimate detail? How are atoms and the ways that they combine related to the properties of the matter that we can measure, such as color, hardness, chemical reactivity, and electrical conductivity?

What fundamental factors influence the stability of a substance? How can we force a desired (but energetically unfavorable) change to take place? What factors control the rate at which a chemical change takes place?

In your study of chemistry, you will learn about these and many other basic ideas that chemists have developed to help them describe and understand the behavior of matter. Along the way, we hope that you come to appreciate the development of this science, one of the grandest intellectual achievements of human endeavor. You will also learn how to apply these fundamental principles to solve real problems. One of your major goals in the study of chemistry should be to develop your ability to think critically and to solve problems (not just do numerical calculations!). In other words, you need to learn to manipulate not only numbers, but also quantitative ideas, words, and concepts.

In the first chapter, our main goals are (1) to begin to get an idea of what chemistry is about and the ways in which chemists view and describe the material world and (2) to acquire some skills that are useful and necessary in the understanding of chemistry, its contribution to science and engineering, and its role in our daily lives.

1-1 MATTER AND ENERGY

Matter is anything that has mass and occupies space. Mass is a measure of the quantity of matter in a sample of any material. The more massive an object is, the more force is required to put it in motion. All bodies consist of matter. Our senses of sight and touch usually tell us that an object occupies space. In the case of colorless, odorless, tasteless gases (such as air), our senses may fail us.

We might say that we can "touch" air when it blows in our faces, but we depend on other evidence to show that a still body of air fits our definition of matter.

Energy is defined as the capacity to do work or to transfer heat. We are familiar with many forms of energy, including mechanical energy, light energy, electrical energy, and heat energy. Light energy from the sun is used by plants as they grow, electrical energy allows us to light a room by flicking a switch, and heat energy cooks our food and warms our homes. Energy can be classified into two principal types: kinetic energy and potential energy.

The term comes from the Greek word *kinein*, meaning "to move." The word "cinema" is derived from the same Greek word.

A body in motion, such as a rolling boulder, possesses energy because of its motion. Such energy is called **kinetic energy.** Kinetic energy represents the capacity for doing work directly. It is easily transferred between objects. **Potential energy** is the energy an object possesses because of its position, condition, or composition. Coal, for example, possesses chemical energy, a form of potential energy, because of its composition. Many electrical generating plants burn coal, producing heat, which is converted to electrical energy. A boulder located atop a mountain possesses potential energy because of its height. It can roll down the mountainside and convert its potential energy into kinetic energy. We discuss energy because all chemical processes are accompanied by energy changes. As some processes occur, energy is released to the surroundings, usually as heat energy. We call such processes **exothermic.** Any combustion (burning) reaction is exothermic. Some chemical reactions and physical changes, however, are **endothermic;** that is, they absorb energy from their surroundings. An example of a physical change that is endothermic is the melting of ice.

Nuclear energy is an important kind of potential energy.

The Law of Conservation of Matter

When we burn a sample of metallic magnesium in the air, the magnesium combines with oxygen from the air (Figure 1-1) to form magnesium oxide, a white powder. This chemical reaction is accompanied by the release of large amounts of heat energy and light energy. When we weigh the product of the reaction, magnesium oxide, we find that it is heavier than the original piece of magnesium. The increase in the mass of a solid is due to the combination of oxygen from the air with magnesium to form magnesium oxide. Many experiments have shown that the mass of the magnesium oxide is exactly the sum of the masses of magnesium and oxygen that combined to form it. Similar statements can be made for all chemical reactions. These observations are summarized in the **Law of Conservation of Matter:**

Figure 1-1 Magnesium burns in the oxygen of the air to form magnesium oxide, a white solid. The mass of magnesium oxide formed is equal to the sum of the masses of oxygen and magnesium that formed it.

> There is no observable change in the quantity of matter during a chemical reaction or during a physical change.

This statement is an example of a **scientific (natural) law,** a general statement based on the observed behavior of matter to which no exceptions are known. A nuclear reaction is *not* a chemical reaction.

The Law of Conservation of Energy

In exothermic chemical reactions, *chemical energy* is usually converted into *heat energy.* Some exothermic processes involve other kinds of energy changes. For example, some liberate light energy without heat, and others produce electrical energy without heat or light. In *endothermic* reactions, heat energy, light energy, or electrical energy is converted into chemical energy. Although chemical changes always involve energy changes, some energy transformations do not involve chemical changes at all. For example, heat energy may be converted into electrical energy or into mechanical energy without any simultaneous chemical changes. Many experiments have demonstrated that all of the energy involved in any chemical or physical change appears in some form after the change. These observations are summarized in the **Law of Conservation of Energy:**

Electricity is produced in hydroelectric plants by the conversion of mechanical energy (from flowing water) into electrical energy.

> Energy cannot be created or destroyed in a chemical reaction or in a physical change. It can only be converted from one form to another.

The Law of Conservation of Matter and Energy

With the dawn of the nuclear age in the 1940s, scientists, and then the world, became aware that matter can be converted into energy. In nuclear reactions (Chapter 26), matter is transformed into energy. The relationship between matter and energy is given by Albert Einstein's now famous equation

Einstein formulated this equation in 1905 as a part of his theory of relativity. Its validity was demonstrated in 1939 with the first controlled nuclear reaction.

$$E = mc^2$$

This equation tells us that the amount of energy released when matter is transformed into energy is the product of the mass of matter transformed and the speed of light squared. At the present time, we have not (knowingly) observed the transformation of energy into matter on a large scale. It does, however, happen on an extremely small scale in "atom smashers," or particle accelerators, used to induce nuclear reactions. Now that the equivalence of matter and energy is recognized, the **Law of Conservation of Matter and Energy** can be stated in a single sentence:

> The combined amount of matter and energy in the universe is fixed.

1-2 STATES OF MATTER

See the *Saunders Interactive General Chemistry CD-ROM,* Screen 1.3, States of Mattter.

Matter can be classified into three states (Figure 1-2), although most of us can think of examples that do not fit neatly into any of the three categories. In the **solid** state, substances are rigid and have definite shapes. Volumes of solids do not vary much with changes in temperature and pressure. In many solids, called crystalline solids, the individual particles that make up the solid occupy definite positions in the crystal structure. The strengths of interaction between the individual particles determine how hard and how strong the crystals are. In the **liquid** state, the individual particles are confined to a given volume. A liquid flows and assumes the shape of its container up to the volume of the liquid. Liquids are very hard to compress. **Gases** are much less dense than liquids and solids. They occupy all parts of any vessel in which they are confined. Gases are capable of infinite expansion and are compressed easily. We conclude that they consist primarily of empty space, meaning that the individual particles are quite far apart.

1-3 CHEMICAL AND PHYSICAL PROPERTIES

To distinguish among samples of different kinds of matter, we determine and compare their **properties.** We recognize different kinds of matter by their properties, which are broadly classified into chemical properties and physical properties.

The properties of a person include height, weight, sex, skin and hair color, and the many subtle features that constitute that person's general appearance.

Chemical properties are exhibited by matter as it undergoes changes in composition. These properties of substances are related to the kinds of chemical changes that the substances undergo. For instance, we have already described the combination of metallic magnesium with gaseous oxygen to form magnesium oxide, a white powder. A chemical property of magnesium is that it can combine with oxygen, releasing energy in the process. A chemical property of oxygen is that it can combine with magnesium.

See the *Saunders Interactive General Chemistry CD-ROM,* Screen 1.2, Physical Properties of Matter.

All substances also exhibit **physical properties** that can be observed in the *absence of any change in composition.* Color, density, hardness, melting point, boiling point, and electrical and thermal conductivities are physical properties. Some physical properties of a

Property	Solid	Liquid	Gas
Rigidity	Rigid	Flows and assumes shape of container	Fills any container completely
Expansion on heating	Slight	Slight	Expands infinitely
Compressibility	Slight	Slight	Easily compressed

Figure 1-2 A comparison of some physical properties of the three states of matter. *(Left)* Iodine, a solid element. *(Center)* Bromine, a liquid element. *(Right)* Chlorine, a gaseous element.

substance depend on the conditions, such as temperature and pressure, under which they are measured. For instance, water is a solid (ice) at low temperatures but is a liquid at higher temperatures. At still higher temperatures, it is a gas (steam). As water is converted from one state to another, its composition is constant. Its chemical properties change very little. On the other hand, the physical properties of ice, liquid water, and steam are different (Figure 1-3).

Properties of matter can be further classified according to whether or not they depend on the *amount* of substance present. The volume and the mass of a sample depend on, and are directly proportional to, the amount of matter in that sample. Such properties, which depend on the amount of material examined, are called **extensive properties.** By contrast, the color and the melting point of a substance are the same for a small sample and for a large one. Properties such as these, which are independent of the amount of material examined, are called **intensive properties.** All chemical properties are intensive properties.

www Many compilations of chemical and physical properties of matter can be found on the World Wide Web. One site is maintained by the U. S. National Institute of Standards and Technology (NIST) at **http://webbook.nist.gov** Perhaps you can find other sites.

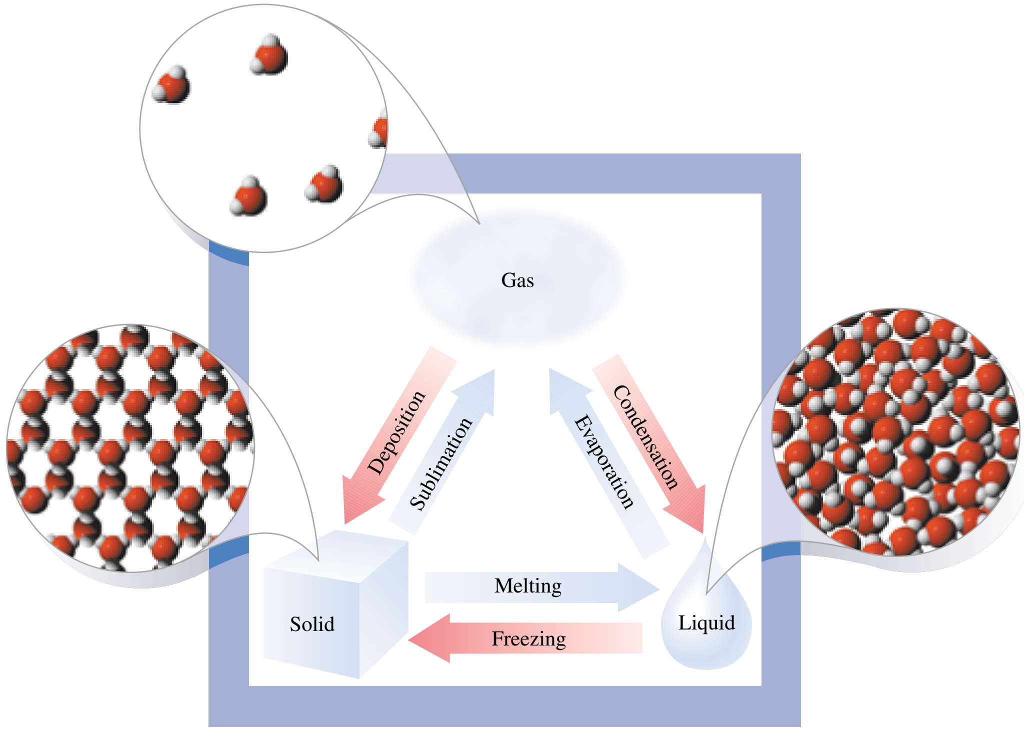
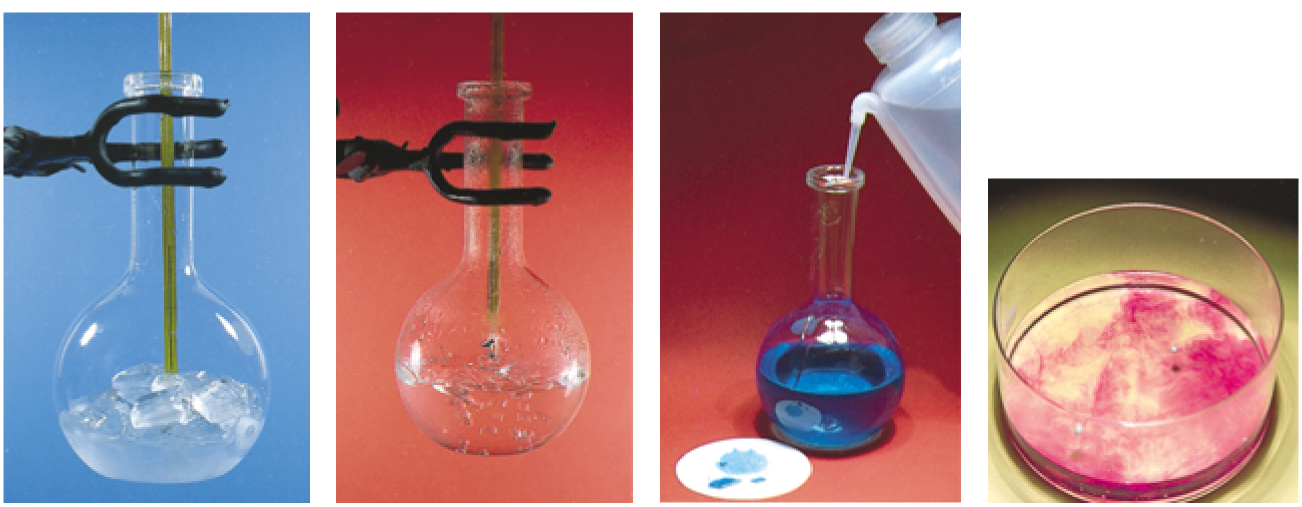


Figure 1-3 Physical changes that occur among the three states of matter. *Sublimation* is the conversion of a solid directly to a gas without passing through the liquid state; the reverse of that process is called *deposition*. The changes shown in blue are endothermic (absorb heat); those shown in red are exothermic (release heat). Water is a substance that is familiar to us in all three physical states. The molecules are close together in the solid and the liquid but far apart in the gas. The molecules in the solid are relatively fixed in position, but those in the liquid and gas can flow around each other.



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

Figure 1-4 Some physical and chemical properties of water. *Physical:* (a) It melts at 0°C; (b) it boils at 100°C (at normal atmospheric pressure); (c) it dissolves a wide range of substances, including copper(II) sulfate, a blue solid. *Chemical:* (d) It reacts with sodium to form hydrogen gas and a solution of sodium hydroxide. The solution contains a little phenolphthalein, which is pink in the presence of sodium hydroxide.

TABLE 1-1 *Physical Properties of a Few Common Substances (at 1 atm pressure)*

			Solubility at 25°C (g/100 g)		
Substance	**Melting Point (°C)**	**Boiling Point (°C)**	***In water***	***In ethyl alcohol***	**Density (g/cm^3)**
acetic acid	16.6	118.1	infinite	infinite	1.05
benzene	5.5	80.1	0.07	infinite	0.879
bromine	−7.1	58.8	3.51	infinite	3.12
iron	1530	3000	insoluble	insoluble	7.86
methane	−182.5	−161.5	0.0022	0.033	6.67×10^{-4}
oxygen	−218.8	−183.0	0.0040	0.037	1.33×10^{-3}
sodium chloride	801	1473	36.5	0.065	2.16
water	0	100	—	infinite	1.00

Because no two different substances have identical sets of chemical and physical properties under the same conditions, we are able to identify and distinguish among different substances. For instance, water is the only clear, colorless liquid that freezes at 0°C, boils at 100°C at one atmosphere of pressure, dissolves a wide variety of substances (e.g., copper(II) sulfate), and reacts violently with sodium (Figure 1-4). Table 1-1 compares several physical properties of a few substances. A sample of any of these substances can be distinguished from the others by observing their properties.

One atmosphere of pressure is the average atmospheric pressure at sea level.

1-4 CHEMICAL AND PHYSICAL CHANGES

We described the reaction of magnesium as it burns in the oxygen of the air (see Figure 1-1). This reaction is a *chemical change.* In any **chemical change,** (1) one or more substances are used up (at least partially), (2) one or more new substances are formed, and (3) energy is absorbed or released. As substances undergo chemical changes, they demonstrate their chemical properties. A **physical change,** on the other hand, occurs with *no change in chemical composition.* Physical properties are usually altered significantly as matter undergoes physical changes (Figure 1-3). In addition, a physical change *may* suggest that a chemical change has also taken place. For instance, a color change, a warming, or the formation of a solid when two solutions are mixed could indicate a chemical change.

See the *Saunders Interactive General Chemistry CD-ROM,* Screens 1.11, Chemical Change, and 1.12, Chemical Change on the Molecular Scale.

Energy is always released or absorbed when chemical or physical changes occur. Energy is required to melt ice, and energy is required to boil water. Conversely, the condensation of steam to form liquid water always liberates energy, as does the freezing of liquid

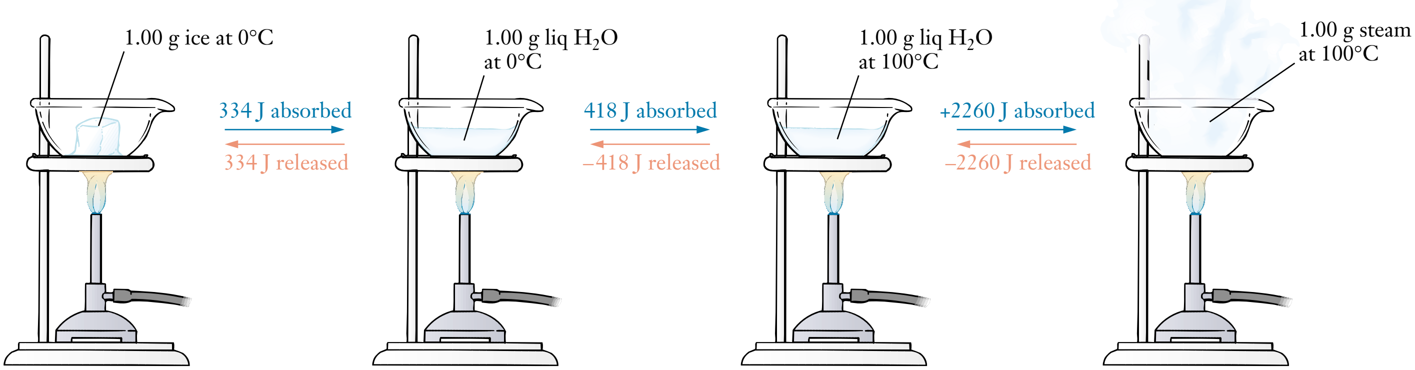


Figure 1-5 Changes in energy that accompany some physical changes for water. The energy unit joules (J) is defined in Section 1-13.

water to form ice. The changes in energy that accompany these physical changes for water are shown in Figure 1-5. At a pressure of one atmosphere, ice always melts at the same temperature (0°C), and pure water always boils at the same temperature (100°C).

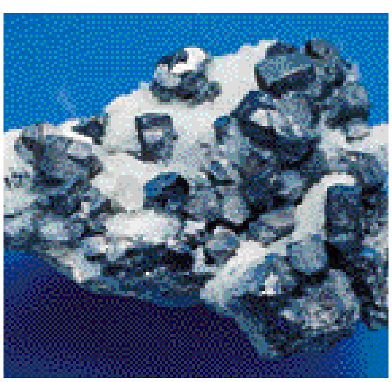
1-5 MIXTURES, SUBSTANCES, COMPOUNDS, AND ELEMENTS

By "composition of a mixture," we mean both the identities of the substances present and their relative amounts in the mixture.

The blue copper(II) sulfate solution in Figure 1-4c is a homogeneous mixture.

Mixtures are combinations of two or more pure substances in which each substance retains its own composition and properties. Almost every sample of matter that we ordinarily encounter is a mixture. The most easily recognized type of mixture is one in which different portions of the sample have recognizably different properties. Such a mixture, which is not uniform throughout, is called **heterogeneous.** Examples include mixtures of salt and charcoal (in which two components with different colors can be distinguished readily from each other by sight), foggy air (which includes a suspended mist of water droplets), and vegetable soup. Another kind of mixture has uniform properties throughout; such a mixture is described as a **homogeneous mixture** and is also called a **solution.** Examples include salt water; some *alloys*, which are homogeneous mixtures of metals in the solid state; and air (free of particulate matter or mists). Air is a mixture of gases. It is mainly nitrogen, oxygen, argon, carbon dioxide, and water vapor. There are only trace amounts of other substances in the atmosphere.

An important characteristic of all mixtures is that they can have variable composition. (For instance, we can make an infinite number of different mixtures of salt and sugar by varying the relative amounts of the two components used.) Consequently, repeating the same experiment on mixtures from different sources may give different results, whereas the same treatment of a pure sample will always give the same results. When the distinction between homogeneous mixtures and pure substances was realized and methods were developed (in the late 1700s) for separating mixtures and studying pure substances, consistent results could be obtained. This resulted in reproducible chemical properties, which formed the basis of real progress in the development of chemical theory.



A heterogeneous mixture of two minerals: galena (black) and quartz (white).

The Resources of the Ocean

As is apparent to anyone who has swum in the ocean, sea water is not pure water but contains a large amount of dissolved solids. In fact, each cubic kilometer of seawater contains about 3.6×10^{10} kilograms of dissolved solids. Nearly 71% of the earth's surface is covered with water. The oceans cover an area of 361 million square kilometers at an average depth of 3729 meters, and hold approximately 1.35 billion cubic kilometers of water. This means that the oceans contain a total of more than 4.8×10^{21} kilograms of dissolved material (or more than 100,000,000,000,000,000,000 pounds). Rivers flowing into the oceans and submarine volcanoes constantly add to this storehouse of minerals. The formation of sediment and the biological demands of organisms constantly remove a similar amount.

Sea water is a very complicated solution of many substances. The main dissolved component of sea water is sodium chloride, common salt. Besides sodium and chlorine, the main elements in sea water are magnesium, sulfur, calcium, potassium, bromine, carbon, nitrogen, and strontium. Together these 10 elements make up more than 99% of the dissolved materials in the oceans. In addition to sodium chloride, they combine to form such compounds as magnesium chloride, potassium sulfate, and calcium carbonate (lime). Animals absorb the latter from the sea and build it into bones and shells.

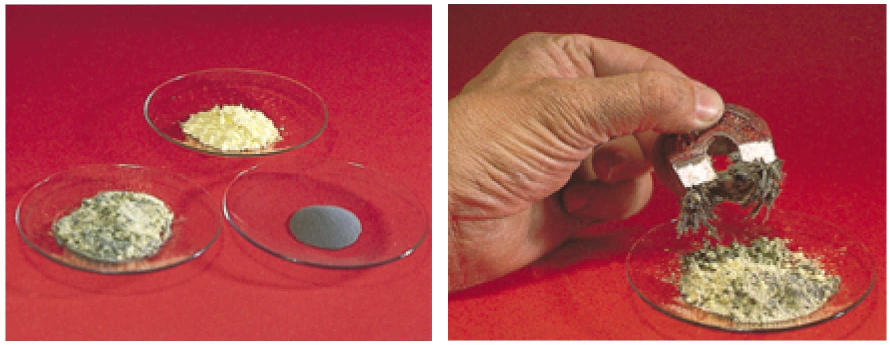
Many other substances exist in smaller amounts in sea water. In fact, most of the 92 naturally occurring elements have been measured or detected in sea water, and the remainder will probably be found as more sensitive analytical techniques become available. There are staggering amounts of valuable metals in sea water, including approximately 1.3×10^{11} kilograms of copper, 4.2×10^{12} kilograms of uranium, 5.3×10^{9} kilograms of gold, 2.6×10^{9} kilograms of silver, and 6.6×10^{8} kilograms of lead. Other elements of economic importance include 2.6×10^{12} kilograms of aluminum, 1.3×10^{10} kilograms of tin, 26×10^{11} kilograms of manganese, and 4.0×10^{10} kilograms of mercury.

One would think that with such a large reservoir of dissolved solids, considerabe "chemical mining" of the ocean would occur. At present, only four elements are commercially extracted in large quantities. They are sodium and chlorine, which are produced from the sea by solar evaporation; magnesium; and bromine. In fact, most of the U. S. production of magnesium is derived from sea water, and the ocean is one of the principal sources of bromine. Most of the other elements are so thinly scattered through the ocean that the cost of their recovery would be much higher than their economic value. However, it is probable that as resources become more and more depleted from the continents, and as recovery techniques become more efficient, mining of sea water will become a much more desirable and feasible prospect.

One promising method of extracting elements from sea water uses marine organisms. Many marine animals concentrate certain elements in their bodies at levels many times higher than the levels in sea water. Vanadium, for example, is taken up by the mucus of certain tunicates and can be concentrated in these animals to more than 280,000 times its concentration in sea water. Other marine organisms can concentrate copper and zinc by a factor of about 1 million. If these animals could be cultivated in large quantities without endangering the ocean ecosystem, they could become a valuable source of trace metals.

In addition to dissolved materials, sea water holds a great store of suspended particulate matter that floats through the water. Some 15% of the manganese in sea water is present in particulate form, as are appreciable amounts of lead and iron. Similarly, most of the gold in sea water is thought to adhere to the surfaces of clay minerals in suspension. As in the case of dissolved solids, the economics of filtering these very fine particles from sea water are not favorable at present. However, because many of the particles suspended in sea water carry an electric charge, ion exchange techniques and modifications of electrostatic processes may someday provide important methods for the recovery of trace metals.

See the *Saunders Interactive General Chemistry CD-ROM,* Screen 1.13, Mixtures and Pure Substances.



(a) **(b)**

Figure 1-6 (a) A mixture of iron and sulfur is a *heterogeneous* mixture. (b) Like any mixture, it can be separated by physical means, such as removing the iron with a magnet.

See the *Saunders Interactive General Chemistry CD-ROM,* Screen 1.14, Separation of Mixtures.

Mixtures can be separated by physical means because each component retains its properties (Figures 1-6 and 1-7). For example, a mixture of salt and water can be separated by evaporating the water and leaving the solid salt behind. To separate a mixture of sand and salt, we could treat it with water to dissolve the salt, collect the sand by filtration, and then evaporate the water to reclaim the solid salt. Very fine iron powder can be mixed with powdered sulfur to give what appears to the naked eye to be a homogeneous mixture of the two. Separation of the components of this mixture is easy, however. The iron may be removed by a magnet, or the sulfur may be dissolved in carbon disulfide, which does not dissolve iron (Figure 1-6).

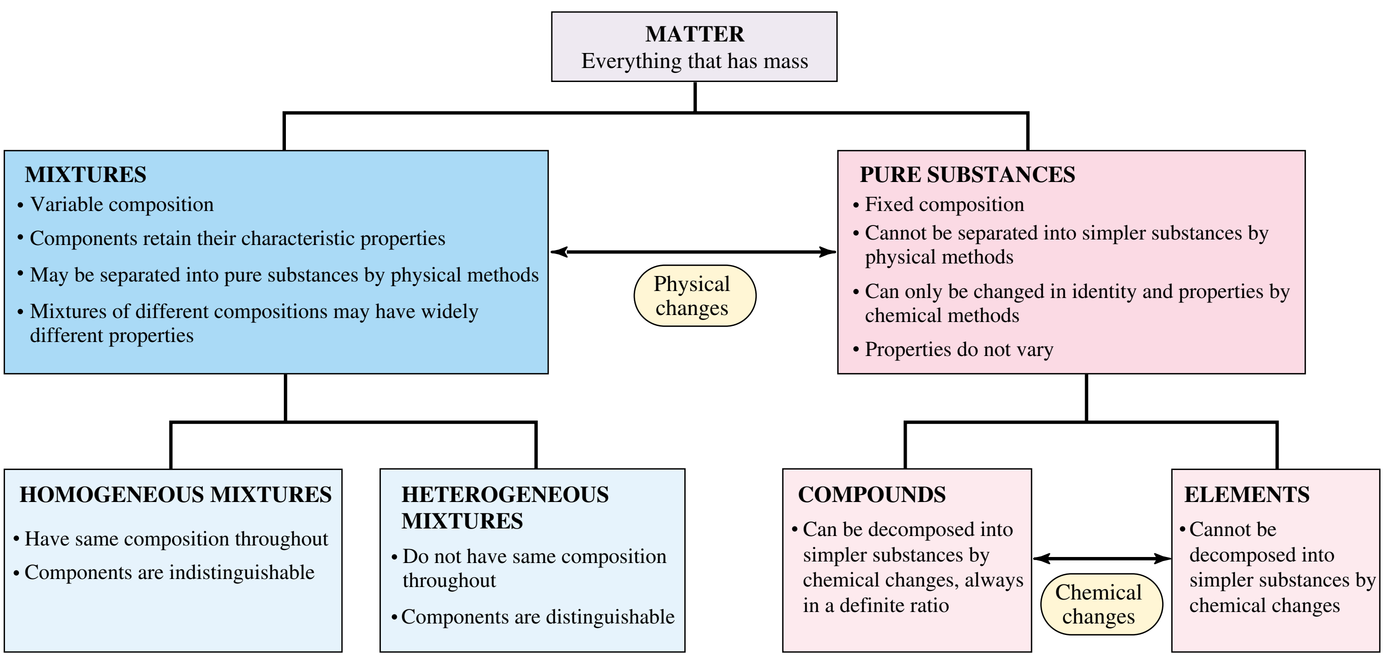


Figure 1-7 One scheme for classification of matter. Arrows indicate the general means by which matter can be separated.

In *any* mixture, (1) the composition can be varied and (2) each component of the mixture retains its own properties.

Imagine that we have a sample of muddy river water (a heterogeneous mixture). We might first separate the suspended dirt from the liquid by filtration. Then we could remove dissolved air by warming the water. Dissolved solids might be removed by cooling the sample until some of it freezes, pouring off the liquid, and then melting the ice. Other dissolved components might be separated by distillation or other methods. Eventually we would obtain a sample of pure water that could not be further separated by any physical separation methods. No matter what the original source of the impure water—the ocean, the Mississippi River, a can of tomato juice, and so on—water samples obtained by purification all have identical composition, and, under identical conditions, they all have identical properties. Any such sample is called a substance, or sometimes a pure substance.

The first ice that forms is quite pure. The dissolved solids tend to stay behind in the remaining liquid.

A **substance** cannot be further broken down or purified by physical means. A substance is matter of a particular kind. Each substance has its own characteristic properties that are different from the set of properties of any other substance.

If we use the definition given here of a *substance*, the phrase *pure substance* may appear to be redundant.

Now suppose we decompose some water by passing electricity through it (Figure 1-8). (An *electrolysis* process is a chemical reaction.) We find that the water is converted into two simpler substances, hydrogen and oxygen; more significantly, hydrogen and

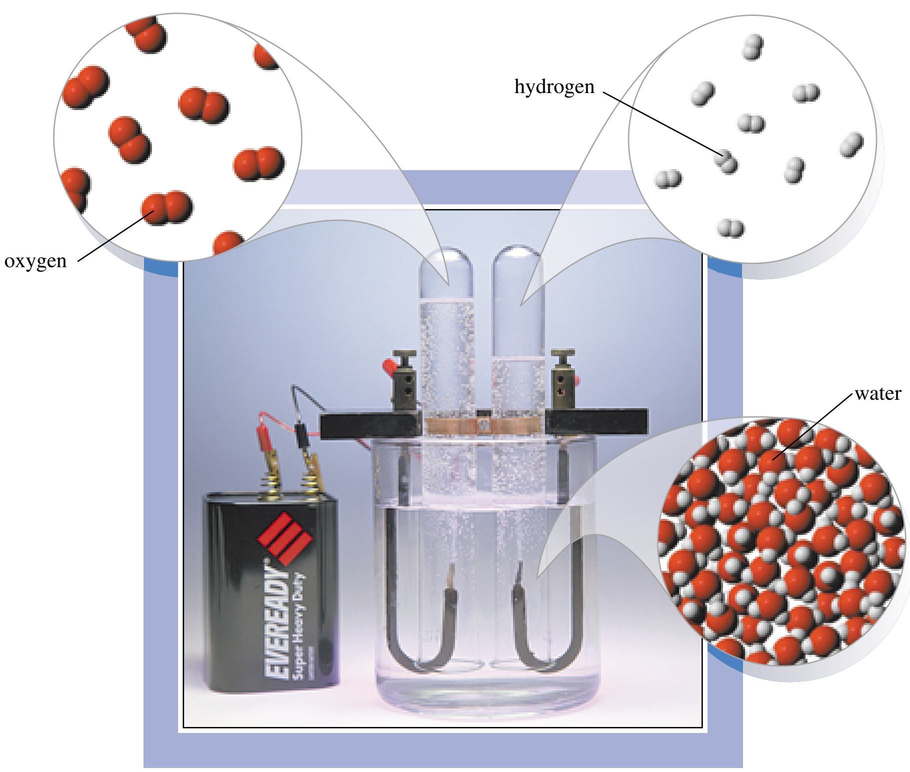


Figure 1-8 Electrolysis apparatus for small-scale chemical decomposition of water by electrical energy. The volume of hydrogen produced (*right*) is twice that of oxygen (*left*). Some dilute sulfuric acid is added to increase the conductivity.

oxygen are *always* present in the same ratio by mass, 11.1% to 88.9%. These observations allow us to identify water as a compound.

A **compound** is a substance that can be decomposed by chemical means into simpler substances, always in the same ratio by mass.

As we continue this process, starting with any substance, we eventually reach a stage at which the new substances formed cannot be further broken down by chemical means. The substances at the end of this chain are called elements.

An **element** is a substance that cannot be decomposed into simpler substances by chemical changes.

For instance, neither of the two gases obtained by the electrolysis of water—hydrogen and oxygen—can be further decomposed, so we know that they are elements.

As another illustration (Figure 1-9), pure calcium carbonate (a white solid present in limestone and seashells) can be broken down by heating to give another white solid (call it A) and a gas (call it B) in the mass ratio 56.0 : 44.0. This observation tells us that calcium carbonate is a compound. The white solid A obtained from calcium carbonate can be further broken down into a solid and a gas in a definite ratio by mass, 71.5 : 28.5. But neither of these can be further decomposed, so they must be elements. The gas is identical to the oxygen obtained from the electrolysis of water; the solid is a metallic element called calcium. Similarly, the gas B, originally obtained from calcium carbonate, can be decomposed into two elements, carbon and oxygen, in a fixed mass ratio, 27.3 : 72.7. This sequence illustrates that a compound can be broken apart into simpler substances at a fixed mass ratio; those simpler substances may be either elements or simpler compounds.

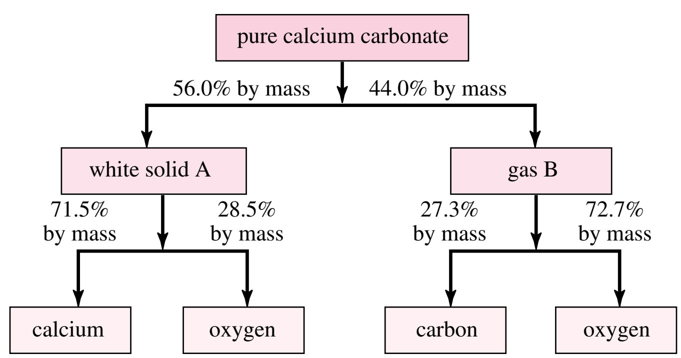


Figure 1-9 Diagram of the decomposition of calcium carbonate to give a white solid A (56.0% by mass) and a gas B (44.0% by mass). This decomposition into simpler substances at a fixed ratio proves that calcium carbonate is a compound. The white solid A further decomposes to give the elements calcium (71.5% by mass) and oxygen (28.5% by mass). This proves that the white solid A is a compound; it is known as calcium oxide. The gas B also can be broken down to give the elements carbon (27.3% by mass) and oxygen (72.7% by mass). This establishes that gas B is a compound; it is known as carbon dioxide.

Furthermore, we may say that *a compound is a pure substance consisting of two or more different elements in a fixed ratio.* Water is 11.1% hydrogen and 88.9% oxygen by mass. Similarly, carbon dioxide is 27.3% carbon and 72.7% oxygen by mass, and calcium oxide (the white solid A in the previous discussion) is 71.5% calcium and 28.5% oxygen by mass. We could also combine the numbers in the previous paragraph to show that calcium carbonate is 40.1% calcium, 12.0% carbon, and 47.9% oxygen by mass. Observations such as these on innumerable pure compounds led to the statement of the **Law of Definite Proportions** (also known as the **Law of Constant Composition**):

> Different samples of any pure compound contain the same elements in the same proportions by mass.

Figure 1-10 The reaction of sodium, a solid element, and chlorine, a gaseous element, to produce sodium chloride (table salt). This reaction gives off considerable energy in the form of heat and light.

The physical and chemical properties of a compound are different from the properties of its constituent elements. Sodium chloride is a white solid that we ordinarily use as table salt (Figure 1-10). This compound is formed by the combination of the element sodium (a soft, silvery white metal that reacts violently with water; see Figure 1-4d) and the element chlorine (a pale green, corrosive, poisonous gas; see Figure 1-2c).

Recall that elements are substances that cannot be decomposed into simpler substances by chemical changes. Nitrogen, silver, aluminum, copper, gold, and sulfur are other examples of elements.

We use a set of **symbols** to represent the elements. These symbols can be written more quickly than names, and they occupy less space. The symbols for the first 109 elements consist of either a capital letter *or* a capital letter and a lowercase letter, such as C (carbon) or Ca (calcium). A list of the known elements and their symbols is given inside the front cover.

In the past, the discoverers of elements claimed the right to name them (see the essay "The Names of the Elements" on page 68), although the question of who had actually discovered the elements first was sometimes disputed. In modern times, new elements are given temporary names and three-letter symbols based on a numerical system. These designations are used until the question of the right to name the newly discovered elements is resolved. Decisions resolving the names of elements 104 through 109 were announced in 1997 by the International Union of Pure and Applied Chemistry (IUPAC), an international organization that represents chemical societies from 40 countries. IUPAC makes recommendations regarding many matters of convention and terminology in chemistry. These recommendations carry no legal force, but they are normally viewed as authoritative throughout the world.

A short list of symbols of common elements is given in Table 1-2. Learning this list will be helpful. Many symbols consist of the first one or two letters of the element's English name. Some are derived from the element's Latin name (indicated in parentheses in Table 1-2) and one, W for tungsten, is from the German *Wolfram.* Names and symbols for additional elements should be learned as they are encountered.

Most of the earth's crust is made up of a relatively small number of elements. Only 10 of the 88 naturally occurring elements make up more than 99% by mass of the earth's crust, oceans, and atmosphere (Table 1-3). Oxygen accounts for roughly half. Relatively few elements, approximately one fourth of the naturally occurring ones, occur in nature as free elements. The rest are always found chemically combined with other elements.

The other known elements have been made artificially in laboratories, as described in Chapter 26.

A very small amount of the matter in the earth's crust, oceans, and atmosphere is involved in living matter. The main element in living matter is carbon, but only a tiny

Mercury is the only metal that is a liquid at room temperature.

The stable form of sulfur at room temperature is a solid.

TABLE 1-2 *Some Common Elements and Their Symbols*

Symbol	Element	Symbol	Element	Symbol	Element
Ag	silver (*argentum*)	F	fluorine	Ni	nickel
Al	aluminum	Fe	iron (*ferrum*)	O	oxygen
Au	gold (*aurum*)	H	hydrogen	P	phosphorus
B	boron	He	helium	Pb	lead (*plumbum*)
Ba	barium	Hg	mercury (*hydrargyrum*)	Pt	platinum
Bi	bismuth	I	iodine	S	sulfur
Br	bromine	K	potassium (*kalium*)	Sb	antimony (*stibium*)
C	carbon	Kr	krypton	Si	silicon
Ca	calcium	Li	lithium	Sn	tin (*stannum*)
Cd	cadmium	Mg	magnesium	Sr	strontium
Cl	chlorine	Mn	manganese	Ti	titanium
Co	cobalt	N	nitrogen	U	uranium
Cr	chromium	Na	sodium (*natrium*)	W	tungsten (*Wolfram*)
Cu	copper (*cuprum*)	Ne	neon	Zn	zinc

TABLE 1-3 *Abundance of Elements in the Earth's Crust, Oceans, and Atmosphere*

Element	Symbol	% by Mass		Element	Symbol	% by Mass	
oxygen	O	49.5%	99.2%	chlorine	Cl	0.19%	0.7%
silicon	Si	25.7		phosphorus	P	0.12	
aluminum	Al	7.5		manganese	Mn	0.09	
iron	Fe	4.7		carbon	C	0.08	
calcium	Ca	3.4		sulfur	S	0.06	
sodium	Na	2.6		barium	Ba	0.04	
potassium	K	2.4		chromium	Cr	0.033	
magnesium	Mg	1.9		nitrogen	N	0.030	
hydrogen	H	0.87		fluorine	F	0.027	
titanium	Ti	0.58		zirconium	Zr	0.023	
				All others combined			≈0.1%

fraction of the carbon in the environment occurs in living organisms. More than one quarter of the total mass of the earth's crust, oceans, and atmosphere is made up of silicon, yet it has almost no biological role.

1-6 MEASUREMENTS IN CHEMISTRY

In the next section, we introduce the standards for basic units of measurement. These standards were selected because they allow us to make precise measurements and because they are reproducible and unchanging. The values of fundamental units are arbitrary.[1] In

[1] *Prior to the establishment of the National Bureau of Standards in 1901, at least 50 different distances had been used as "1 foot" in measuring land within New York City. Thus the size of a 100-ft by 200-ft lot in New York City depended on the generosity of the seller and did not necessarily represent the expected dimensions.*

TABLE 1-4 *The Seven Fundamental Units of Measurement (SI)*

Physical Property	Name of Unit	Symbol
length	meter	m
mass	kilogram	kg
time	second	s
electric current	ampere	A
temperature	kelvin	K
luminous intensity	candela	cd
amount of substance	mole	mol

the United States, all units of measure are set by the National Institute of Standards and Technology, NIST (formerly the National Bureau of Standards, NBS). Measurements in the scientific world are usually expressed in the units of the metric system or its modernized successor, the International System of Units (SI). The SI, adopted by the National Bureau of Standards in 1964, is based on the seven fundamental units listed in Table 1-4. All other units of measurement are derived from them.

The abbreviation SI comes from the French *le Système International.*

In this text we shall use both metric units and SI units. Conversions between non-SI and SI units are usually straightforward. Appendix C lists some important units of measurement and their relationships to one another. Appendix D lists several useful physical constants. The most frequently used of these appear on the inside back cover.

See the *Saunders Interactive General Chemistry CD-ROM,* Screen 1.16, The Metric System.

The metric and SI systems are *decimal systems, in which prefixes are used to indicate fractions and multiples of ten.* The same prefixes are used with all units of measurement. The distances and masses in Table 1-5 illustrate the use of some common prefixes and the relationships among them.

The prefixes used in the SI and metric systems may be thought of as *multipliers.* For example, the prefix *kilo-* indicates multiplication by 1000 or 10^3, and *milli-* indicates multiplication by 0.001 or 10^{-3}.

TABLE 1-5 *Common Prefixes Used in the SI and Metric Systems*

Prefix	Abbreviation	Meaning	Example
mega-	M	10^6	1 megameter (Mm) = 1×10^6 m
kilo-*	k	10^3	1 kilometer (km) = 1×10^3 m
deci-	d	10^{-1}	1 decimeter (dm) = 1×10^{-1} m
centi-*	c	10^{-2}	1 centimeter (cm) = 1×10^{-2} m
milli-*	m	10^{-3}	1 milligram (mg) = 1×10^{-3} g
micro-*	μ†	10^{-6}	1 microgram (μg) = 1×10^{-6} g
nano-*	n	10^{-9}	1 nanogram (ng) = 1×10^{-9} g
pico-	p	10^{-12}	1 picogram (pg) = 1×10^{-12} g

**These prefixes are commonly used in chemistry.*

†This is the Greek letter μ (pronounced "mew").

1-7 UNITS OF MEASUREMENT

Mass and Weight

We distinguish between mass and weight. **Mass** is the measure of the quantity of matter a body contains (see Section 1-1). The mass of a body does not vary as its position changes. On the other hand, the **weight** of a body is a measure of the gravitational attraction of the earth for the body, and this varies with distance from the center of the earth. An object weighs very slightly less high up on a mountain than at the bottom of a deep valley. Because the mass of a body does not vary with its position, the mass of a body is a more fundamental property than its weight. We have become accustomed, however, to using the term "weight" when we mean mass, because weighing is one way of measuring mass (Figure 1-11). Because we usually discuss chemical reactions at constant gravity, weight relationships are just as valid as mass relationships. Nevertheless, we should keep in mind that the two are not identical.

The basic unit of mass in the SI system is the **kilogram** (Table 1-6). The kilogram is defined as the mass of a platinum–iridium cylinder stored in a vault in Sèvres, near Paris, France. A 1-lb object has a mass of 0.4536 kg. The basic mass unit in the earlier *metric system* was the gram. A U.S. five-cent coin (a "nickel") has a mass of about 5 g.

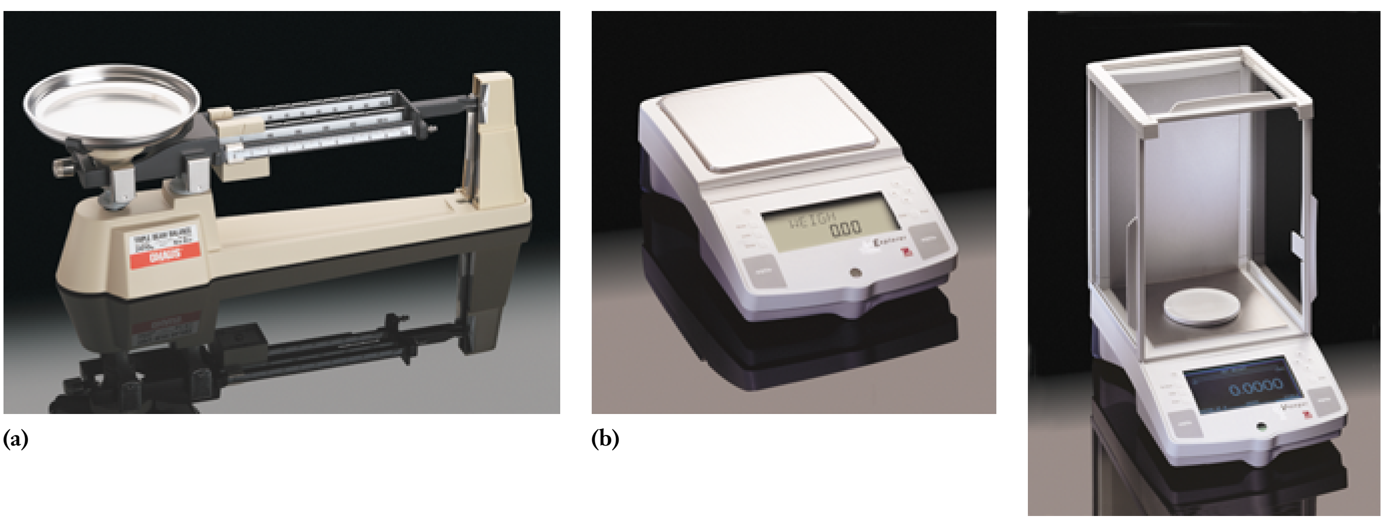
TABLE 1-6 *Some SI Units of Mass*

*kilo*gram, kg	base unit
gram, g	1,000 g = 1 kg
*milli*gram, mg	1,000 mg = 1 g
*micro*gram, μg	1,000,000 μg = 1 g

Length

The meter was originally defined (1791) as one ten-millionth of the distance between the North Pole and the equator.

The **meter** is the standard unit of length (distance) in both SI and metric systems. The meter is defined as the distance light travels in a vacuum in 1/299,792,468 second. It is approximately 39.37 inches. In situations in which the English system would use inches, the metric centimeter (1/100 meter) is convenient. The relationship between inches and centimeters is shown in Figure 1-12.



(a) (b) (c)

Figure 1-11 Three types of laboratory balances. (a) A triple-beam balance used for determining mass to about ±0.01 g. (b) A modern electronic top-loading balance that gives a direct readout of mass to ±0.001 g. (c) A modern analytical balance that can be used to determine mass to ±0.0001 g. Analytical balances are used when masses must be determined as precisely as possible.

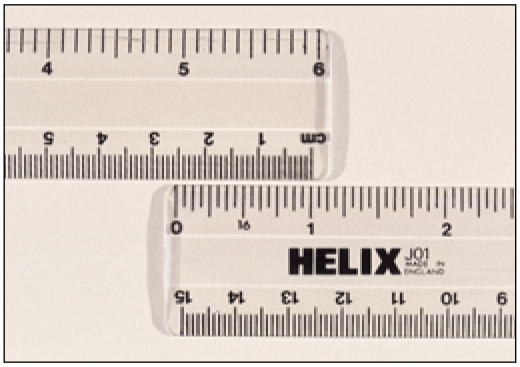


Figure 1-12 The relationship between inches and centimeters: 1 in. = 2.54 cm (exactly).

Figure 1-13 Some laboratory apparatus used to measure volumes of liquids: 150-mL beaker (*bottom left*, green liquid); 25-mL buret (*top left*, red); 1000-mL volumetric flask (*center*, yellow); 100-mL graduated cylinder (*right front*, blue); and 10-mL volumetric pipet (*right rear*, green).

Volume

Volumes are often measured in liters or milliliters in the metric system. One liter (1 L) is one cubic decimeter (1 dm^3), or 1000 cubic centimeters (1000 cm^3). One milliliter (1 mL) is 1 cm^3. In medical laboratories, the cubic centimeter (cm^3) is often abbreviated cc. In the SI, the cubic meter is the basic volume unit and the cubic decimeter replaces the metric unit, liter. Different kinds of glassware are used to measure the volume of liquids. The one we choose depends on the accuracy we desire. For example, the volume of a liquid dispensed can be measured more accurately with a buret than with a small graduated cylinder (Figure 1-13). Equivalences between common English units and metric units are summarized in Table 1-7.

Sometimes we must combine two or more units to describe a quantity. For instance, we might express the speed of a car as 60 mi/h (also mph). Recall that the algebraic notation x^{-1} means $1/x$; applying this notation to units, we see that h^{-1} means 1/h, or "per hour." So the unit of speed could also be expressed as $mi \cdot h^{-1}$.

TABLE 1-7 ***Conversion Factors Relating Length, Volume, and Mass (weight) Units***

	Metric		**English**		**Metric–English Equivalents**	
Length	1 km	$= 10^3$ m	1 ft	= 12 in.	2.54 cm	= 1 in.
	1 cm	$= 10^{-2}$ m	1 yd	= 3 ft	39.37 in.*	= 1 m
	1 mm	$= 10^{-3}$ m	1 mile	= 5280 ft	1.609 km*	= 1 mile
	1 nm	$= 10^{-9}$ m				
	1 Å	$= 10^{-10}$ m				
Volume	1 mL	$= 1$ $cm^3 = 10^{-3}$ L	1 gal	= 4 qt = 8 pt	1L	= 1.057 qt*
	1 m^3	$= 10^6$ $cm^3 = 10^3$ L	1 qt	= 57.75 $in.^3$*	28.32 L	= 1 ft^3
Mass	1 kg	$= 10^3$ g	1 lb	= 16 oz	453.6 g*	= 1 lb
	1 mg	$= 10^{-3}$ g			1 g	= 0.03527 oz*
	1 metric tonne	$= 10^3$ kg	1 short ton	= 2000 lb	1 metric tonne	= 1.102 short ton*

**These conversion factors, unlike the others listed, are inexact. They are quoted to four significant figures, which is ordinarily more than sufficient.*

1-8 USE OF NUMBERS

See the *Saunders Interactive General Chemistry CD-ROM,* Screen 1.17, Using Numerical Information.

In chemistry, we measure and calculate many things, so we must be sure we understand how to use numbers. In this section we discuss two aspects of the use of numbers: (1) the notation of very large and very small numbers and (2) an indication of how well we actually know the numbers we are using. You will carry out many calculations with calculators. Please refer to Appendix A for some instructions about the use of electronic calculators.

Scientific Notation

In exponential form, these numbers are

$$6.02 \times 10^{23} \text{ gold atoms}$$

$$3.27 \times 10^{-22} \text{ gram}$$

We use **scientific notation** when we deal with very large and very small numbers. For example, 197 grams of gold contains approximately

$$602{,}000{,}000{,}000{,}000{,}000{,}000{,}000 \text{ gold atoms}$$

The mass of one gold atom is approximately

$$0.000\ 000\ 000\ 000\ 000\ 000\ 000\ 327 \text{ gram}$$

In using such large and small numbers, it is inconvenient to write down all the zeroes. In scientific (exponential) notation, we place one nonzero digit to the left of the decimal.

$$4{,}300{,}000. = 4.3 \times 10^{6}$$

6 places to the left, ∴ exponent of 10 is 6

$$0.000348 = 3.48 \times 10^{-4}$$

4 places to the right, ∴ exponent of 10 is −4

The reverse process converts numbers from exponential to decimal form. See Appendix A for more detail, if necessary.

Problem-Solving Tip: *Know How to Use Your Calculator*

Students sometimes make mistakes when they try to enter numbers into their calculators in scientific notation. Suppose you want to enter the number 4.36×10^{-2}. On most calculators, you would

(1) press 4.36

(2) press EE or EXP, which stands for “times ten to the”

(3) press 2 (the magnitude of the exponent) and then ± or CHS (to change its sign)

The calculator display might show the value as $\boxed{4.36 \quad -02}$ or as $\boxed{0.0436}$. Different calculators show different numbers of digits, which can sometimes be adjusted.

If you wished to enter -4.36×10^{2}, you would

(1) press 4.36, then press ± or CHS to change its sign,

(2) press EE or EXP, and then press 2

The calculator would then show $\boxed{-4.36 \quad 02}$ or $\boxed{-436.0}$.

Caution: Be sure you remember that the EE or EXP button *includes* the "times 10" operation. An error that beginners often make is to enter " $\times$ 10" explicitly when trying to enter a number in scientific notation. Suppose you mistakenly enter 3.7×10^2 as follows:

(1) enter 3.7

(2) press $\times$ and then enter 10

(3) press EXP or EE and then enter 2

The calculator then shows the result as 3.7×10^3 or 3700—why? This sequence is processed by the calculator as follows: Step (1) enters the number 3.7; step (2) multiplies by 10, to give 37; step (3) multiplies this by 10^2, to give 37×10^2 or 3.7×10^3.

Other common errors include changing the sign of the exponent when the intent was to change the sign of the entire number (e.g., -3.48×10^4 entered as 3.48×10^{-4}).

When in doubt, carry out a trial calculation for which you already know the answer. For instance, multiply 300 by 2 by entering the first value as 3.00×10^2 and then multiplying by 2; you know the answer should be 600, and if you get any other answer, you know you have done something wrong. If you cannot find (or understand) the printed instructions for *your calculator*, your instructor or a classmate might be able to help.

Significant Figures

There are two kinds of numbers. *Numbers obtained by counting or from definitions are* **exact numbers.** They are known to be absolutely accurate. For example, the exact number of people in a closed room can be counted, and there is no doubt about the number of people. A dozen eggs is defined as exactly 12 eggs, no more, no fewer (Figure 1-14).

An *exact* number may be thought of as containing an *infinite* number of significant figures.

(a) **(b)**

Figure 1-14 (a) A dozen eggs is exactly 12 eggs. (b) A specific swarm of honeybees contains an *exact* number of live bees (but it would be difficult to count them, and any two *swarms* would be unlikely to contain the same exact number of bees).

There is some uncertainty in all measurements.

Numbers obtained from measurements are not exact. Every measurement involves an estimate. For example, suppose you are asked to measure the length of this page to the nearest 0.1 mm. How do you do it? The smallest divisions (calibration lines) on a meter stick are 1 mm apart (see Figure 1-12). An attempt to measure to 0.1 mm requires estimation. If three different people measure the length of the page to 0.1 mm, will they get the same answer? Probably not. We deal with this problem by using significant figures.

Significant figures indicate the *uncertainty* in measurements.

Significant figures are digits believed to be correct by the person who makes a measurement. We assume that the person is competent to use the measuring device. Suppose one measures a distance with a meter stick and reports the distance as 343.5 mm. What does this number mean? In this person's judgment, the distance is greater than 343.4 mm but less than 343.6 mm, and the best estimate is 343.5 mm. The number 343.5 mm contains four significant figures. The last digit, 5, is a *best estimate* and is therefore doubtful, but it is considered to be a significant figure. In reporting numbers obtained from measurements, *we report one estimated digit, and no more.* Because the person making the measurement is not certain that the 5 is correct, it would be meaningless to report the distance as 343.53 mm.

To see more clearly the part significant figures play in reporting the results of measurements, consider Figure 1-15a. Graduated cylinders are used to measure volumes of liquids when a high degree of accuracy is not necessary. The calibration lines on a 50-mL graduated cylinder represent 1-mL increments. Estimation of the volume of liquid in a 50-mL cylinder to within 0.2 mL ($\frac{1}{5}$ of one calibration increment) with reasonable certainty is possible. We might measure a volume of liquid in such a cylinder and report the volume as 38.6 mL, that is, to three significant figures.

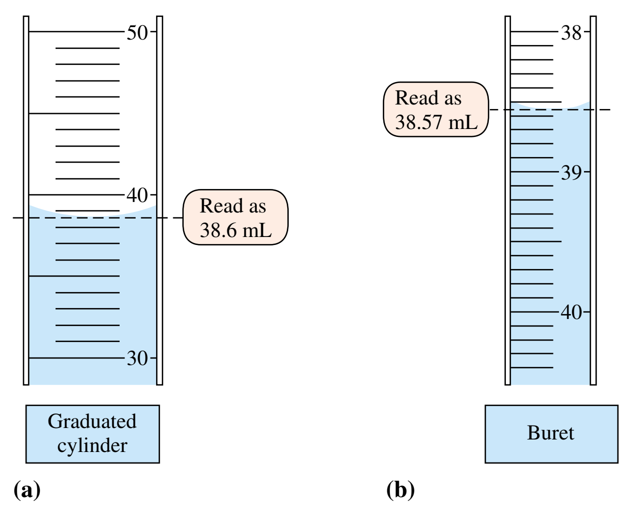


Figure 1-15 Measurement of the volume of water using two types of volumetric glassware. For consistency, we always read the bottom of the meniscus (the curved surface of the water). (a) A graduated cylinder is used to measure the amount of liquid *contained* in the glassware, so the scale increases from bottom to top. The level in a 50-mL graduated cylinder can usually be estimated to within 0.2 mL. The level here is 38.6 mL (three significant figures). (b) We use a buret to measure the amount of liquid *delivered* from the glassware, by taking the difference between an initial and a final volume reading. The level in a 50-mL buret can be read to within 0.02 mL. The level here is 38.57 mL (four significant figures).

Burets are used to measure volumes of liquids when higher accuracy is required. The calibration lines on a 50-mL buret represent 0.1-mL increments, allowing us to make estimates to within 0.02 mL ($\frac{1}{5}$ of one calibration increment) with reasonable certainty (Figure 1-15b). Experienced individuals estimate volumes in 50-mL burets to 0.01 mL with considerable reproducibility. For example, using a 50-mL buret, we can measure out 38.57 mL (four significant figures) of liquid with reasonable accuracy.

Accuracy refers to how closely a measured value agrees with the correct value. **Precision** refers to how closely individual measurements agree with one another. Ideally, all measurements should be both accurate and precise. Measurements may be quite precise yet quite inaccurate because of some *systematic error*, which is an error repeated in each measurement. (A faulty balance, for example, might produce a systematic error.) Very accurate measurements are seldom imprecise.

Measurements are frequently repeated to improve accuracy and precision. Average values obtained from several measurements are usually more reliable than individual measurements. Significant figures indicate how precisely measurements have been made (assuming the person who made the measurements was competent).

Some simple rules govern the use of significant figures.

1. Nonzero digits are always significant.

For example, 38.57 mL has four significant figures; 288 g has three significant figures.

2. Zeroes are sometimes significant, and sometimes they are not.

a. Zeroes at the *beginning* of a number (used just to position the decimal point) are never significant.

For example, 0.052 g has two significant figures; 0.00364 m has three significant figures. These could also be reported in scientific notation (Appendix A) as 5.2×10^{-2} g and 3.64×10^{-3} m, respectively.

b. Zeroes *between* nonzero digits are always significant.

For example, 2007 g has four significant figures; 6.08 km has three significant figures.

c. Zeroes at the *end* of a number that contains a decimal point are always significant.

For example, 38.0 cm has three significant figures; 440.0 m has four significant figures. These could also be reported as 3.80×10^{1} cm and 4.400×10^{2} m, respectively.

d. Zeroes at the *end* of a number that does not contain a decimal point may or may not be significant.

When we wish to specify that all of the zeroes in such a number *are* significant, we may indicate this by placing a decimal point after the number. For instance, 130. grams can represent a mass known to *three* significant figures, that is, 130 ± 1 gram.

For example, the quantity 24,300 km could represent three, four, or five significant figures. We are given insufficient information to answer the question. If both of the zeroes are used just to place the decimal point, the number should appear as 2.43×10^4 km (three significant figures). If only one of the zeroes is used to place the decimal point (i.e., the number was measured ±10), the number is 2.430×10^4 km (four significant figures). If the number is actually known to be 24,300 ± 1, it should be written as 2.4300×10^4 km (five significant figures).

3. Exact numbers can be considered as having an unlimited number of significant figures. This applies to defined quantities.

For example, in the equivalence 1 yard = 3 feet, the numbers 1 and 3 are exact, and we do not apply the rules of significant figures to them. The equivalence 1 inch = 2.54 centimeters is an exact one.

A calculated number can never be more precise than the numbers used to calculate it. The following rules show how to get the number of significant figures in a calculated number.

4. In addition and subtraction, the last digit retained in the sum or difference is determined by the position of the first doubtful digit.

EXAMPLE 1-1 *Significant Figures (Addition and Subtraction)*

(a) Add 37.24 mL and 10.3 mL. (b) Subtract 21.2342 g from 27.87 g.

Plan

We first check to see that the quantities to be added or subtracted are expressed in the same units. We carry out the addition or subtraction. Then we follow Rule 4 for significant figures to express the answer to the correct number of significant figures.

Solution

Doubtful digits are underlined in this example.

(a)

$$\begin{array}{r} 37.2\underline{4}\ \text{mL} \\ +10.\underline{3}\ \ \ \text{mL} \\ \hline 47.\underline{54}\ \text{mL} \end{array}$$

47.54 mL is reported as $\boxed{47.5\text{ mL}}$ (calculator gives 47.54)

(b)

$$\begin{array}{r} 27.8\underline{7}\ \ \ \ \text{g} \\ -21.234\underline{2}\text{g} \\ \hline 6.6\underline{358}\text{g} \end{array}$$

6.6358g is reported as $\boxed{6.64\text{ g}}$ (calculator gives 6.6358)

5. In multiplication and division, an answer contains no more significant figures than the least number of significant figures used in the operation.

EXAMPLE 1-2 *Significant Figures (Multiplication)*

What is the area of a rectangle 1.23 cm wide and 12.34 cm long?

Plan

The area of a rectangle is its length times its width. We must first check to see that the width and length are expressed in the same units. (They are, but if they were not, we must first convert one to the units of the other.) Then we multiply the width by the length. We then follow Rule 5 for significant figures to find the correct number of significant figures. The units for the result are equal to the product of the units for the individual terms in the multiplication.

Solution

$$A = \ell \times w = (12.34\text{ cm})(1.23\text{ cm}) = \boxed{15.2\text{ cm}^2}$$

$$\text{(calculator result} = 15.1782)$$

Because three is the smallest number of significant figures used, the answer should contain only three significant figures. The number generated by an electronic calculator (15.1782) implies more accuracy than is justified; the result cannot be more accurate than the information that led to it. Calculators have no judgment, so you must exercise yours.

You should now work Exercise 27.

With many examples we suggest selected exercises from the end of the chapter. These exercises use the skills or concepts from that example. Now you should work Exercise 27 from the end of this chapter.

The step-by-step calculation in the margin demonstrates why the area is reported as 15.2 cm^2 rather than 15.1782 cm^2. The length, 12.34 cm, contains four significant figures, whereas the width, 1.23 cm, contains only three. If we underline each uncertain figure, as well as each figure obtained from an uncertain figure, the step-by-step multiplication gives the result reported in Example 1-2. We see that there are only two certain figures (15) in the result. We report the first doubtful figure (.2), but no more. Division is just the reverse of multiplication, and the same rules apply.

$$\begin{array}{rl}
12.3\underline{4} & \text{cm} \\
\times\quad 1.2\underline{3} & \text{cm} \\
\hline
\underline{3}\,\underline{7}\,\underline{0}\,\underline{2} & \\
2\,4\,6\,\underline{8} & \\
1\,2\,3\,\underline{4} & \\
\hline
15.\underline{1782} & \text{cm}^2 = 15.2\text{ cm}^2
\end{array}$$

In the three simple arithmetic operations we have performed, the number combination generated by an electronic calculator is not the "answer" in a single case! The correct result of each calculation, however, can be obtained by "rounding off." The rules of significant figures tell us where to round off.

In rounding off, certain conventions have been adopted. When the number to be dropped is less than 5, the preceding number is left unchanged (e.g., 7.34 rounds off to 7.3). When it is more than 5, the preceding number is increased by 1 (e.g., 7.37 rounds off to 7.4). When the number to be dropped is 5, the preceding number is set to the nearest *even* number (e.g., 7.45 rounds off to 7.4, and 7.35 rounds off to 7.4).

Rounding off to an even number is intended to reduce the accumulation of errors in chains of calculations.

✓ Problem-Solving Tip: *When Do We Round?*

When a calculation involves several steps, we often show the answer to each step to the correct number of significant figures. We carry all digits in the calculator to the end of the calculation, however. Then we round the final answer to the appropriate number of significant figures. When carrying out such a calculation, it is safest to carry extra figures through all steps and then to round the final answer appropriately.

1-9 THE UNIT FACTOR METHOD (DIMENSIONAL ANALYSIS)

Many chemical and physical processes can be described by numerical relationships. In fact, many of the most useful ideas in science must be treated mathematically. In this section, we review some problem-solving skills.

It would be nonsense to say that the length of a piece of cloth is 4.7. We must specify units with the number—4.7 inches, 4.7 feet, or 4.7 meters, for instance.

The units must *always* accompany the numeric value of a measurement, whether we are writing about the quantity, talking about it, or using it in calculations.

Multiplication by unity (by one) does not change the value of an expression. If we represent "one" in a useful way, we can do many conversions by just "multiplying by one." This method of performing calculations is known as **dimensional analysis,** the **factor-label method,** or the **unit factor method.** Regardless of the name chosen, it is a powerful mathematical tool that is almost foolproof.

Unit factors may be constructed from any two terms that describe the same or equivalent "amounts" of whatever we may consider. For example, 1 foot is equal to exactly 12 inches, by definition. We may write an equation to describe this equality:

$$1 \text{ ft} = 12 \text{ in.}$$

Dividing both sides of the equation by 1 ft gives

$$\frac{1 \cancel{\text{ft}}}{1 \cancel{\text{ft}}} = \frac{12 \text{ in.}}{1 \text{ ft}} \quad \text{or} \quad 1 = \frac{12 \text{ in.}}{1 \text{ ft}}$$

The factor (fraction) 12 in./1 ft is a unit factor because the numerator and denominator describe the same distance. Dividing both sides of the original equation by 12 in. gives 1 = 1 ft/12 in., a second unit factor that is the reciprocal of the first. *The reciprocal of any unit factor is also a unit factor.* Stated differently, division of an amount by the same amount always yields one!

In the English system we can write many unit factors, such as

Unless otherwise indicated, a "ton" refers to a "short ton," 2000 lb. There are also the "long ton," which is 2240 lb, and the metric tonne, which is 1000 kg.

$$\frac{1 \text{ yd}}{3 \text{ ft}}, \quad \frac{1 \text{ yd}}{36 \text{ in.}}, \quad \frac{1 \text{ mi}}{5280 \text{ ft}}, \quad \frac{4 \text{ qt}}{1 \text{ gal}}, \quad \frac{2000 \text{ lb}}{1 \text{ ton}}$$

The reciprocal of each of these is also a unit factor. Items in retail stores are frequently priced with unit factors, such as 39¢/lb and \$3.98/gal. When all the quantities in a unit factor come from definitions, the unit is known to an unlimited (infinite) number of significant figures. For instance, if you bought eight 1-gallon jugs of something priced at \$3.98/gal, the total cost would be 8 × \$3.98, or \$31.84; the merchant would not round this to \$31.80, let alone to \$30.

In science, nearly all numbers have units. What does 12 mean? Usually we must supply appropriate units, such as 12 eggs or 12 people. In the unit factor method, the units guide us through calculations in a step-by-step process, because all units except those in the desired result cancel.

EXAMPLE 1-3 *Unit Factors*

Express 1.47 mi in inches.

Plan

First we write down the units of what we wish to know, preceded by a question mark. Then we set it equal to whatever we are given:

$$\underline{?}\text{ inches} = 1.47\text{ miles}$$

Then we choose unit factors to convert the given units (miles) to the desired units (inches):

$$\text{miles} \longrightarrow \text{feet} \longrightarrow \text{inches}$$

We relate (a) miles to feet and then (b) feet to inches.

Solution

$$\underline{?}\text{ in.} = 1.47\ \cancel{\text{mi}} \times \frac{5280\ \cancel{\text{ft}}}{1\ \cancel{\text{mi}}} \times \frac{12\text{ in.}}{1\ \cancel{\text{ft}}} = \boxed{9.31 \times 10^4\text{ in.}} \quad \text{(calculator gives 93139.2)}$$

In the interest of clarity, cancellation of units will be omitted in the remainder of this book. You may find it useful to continue the cancellation of units.

Note that both miles and feet cancel, leaving only inches, the desired unit. Thus, there is no ambiguity as to how the unit factors should be written. The answer contains three significant figures because there are three significant figures in 1.47 miles.

Problem-Solving Tip: *Significant Figures*

"*How do defined quantities affect significant figures?*" Any quantity that comes from a *definition* is exact, that is, it is known to an unlimited number of significant figures. In Example 1-3, the quantities 5280 ft, 1 mile, 12 in., and 1 ft all come from definitions, so they do not limit the significant figures in the answer.

Problem-Solving Tip: *Think About Your Answer!*

It is often helpful to ask yourself, "Does the answer make sense?" In Example 1-3, the distance involved is more than a mile. We expect this distance to be many inches, so a large answer is not surprising. Suppose we had mistakenly multiplied by the unit factor $\frac{1\text{ mile}}{5280\text{ feet}}$ (and not noticed that the units did not cancel properly); we would have gotten the answer 3.34×10^{-3} in. (0.00334 in.), which we should have immediately recognized as nonsense!

Within the SI and metric systems, many measurements are related to one another by powers of ten.

EXAMPLE 1-4 *Unit Conversions*

The Ångstrom (Å) is a unit of length, 1×10^{-10} m, that provides a convenient scale on which to express the radii of atoms. Radii of atoms are often expressed in nanometers. The radius of a phosphorus atom is 1.10 Å. What is the distance expressed in centimeters and nanometers?

Å → m → cm

Å → m → nm

Plan

We use the equalities 1 Å = 1×10^{-10} m, 1 cm = 1×10^{-2} m, and 1 nm = 1×10^{-9} m to construct the unit factors that convert 1.10 Å to the desired units.

Solution

$$\underline{?}\ \text{cm} = 1.10\ \text{Å} \times \frac{1 \times 10^{-10}\ \text{m}}{1\ \text{Å}} \times \frac{1\ \text{cm}}{1 \times 10^{-2}\ \text{m}} = \boxed{1.10 \times 10^{-8}\ \text{cm}}$$

$$\underline{?}\ \text{nm} = 1.10\ \text{Å} \times \frac{1.0 \times 10^{-10}\ \text{m}}{1\ \text{Å}} \times \frac{1\ \text{nm}}{1 \times 10^{-9}\ \text{m}} = \boxed{1.10 \times 10^{-1}\ \text{nm}}$$

All the unit factors used in this example contain only exact numbers.

You should now work Exercise 30.

EXAMPLE 1-5 *Volume Calculation*

Assuming a phosphorus atom is spherical, calculate its volume in Å^3, cm^3, and nm^3. The formula for the volume of a sphere is $V = (\frac{4}{3})\pi r^3$. Refer to Example 1-4.

Plan

We use the results of Example 1-4 to calculate the volume in each of the desired units.

Solution

$$\underline{?}\ \text{Å}^3 = (\tfrac{4}{3})\pi(1.10\ \text{Å})^3 = \boxed{5.58\ \text{Å}^3}$$

$$\underline{?}\ \text{cm}^3 = (\tfrac{4}{3})\pi(1.10 \times 10^{-8}\ \text{cm})^3 = \boxed{5.58 \times 10^{-24}\ \text{cm}^3}$$

$$\underline{?}\ \text{nm}^3 = (\tfrac{4}{3})\pi(1.10 \times 10^{-1}\ \text{nm})^3 = \boxed{5.58 \times 10^{-3}\ \text{nm}^3}$$

1 Å = 10^{-10} m = 10^{-8} cm

You should now work Exercise 34.

EXAMPLE 1-6 *Mass Conversion*

A sample of gold has a mass of 0.234 mg. What is its mass in g? in kg?

Plan

We use the relationships 1 g = 1000 mg and 1 kg = 1000 g to write the required unit factors.

Solution

$$\underline{?}\ \text{g} = 0.234\ \text{mg} \times \frac{1\ \text{g}}{1000\ \text{mg}} = \boxed{2.34 \times 10^{-4}\ \text{g}}$$

$$\underline{?}\ \text{kg} = 2.34 \times 10^{-4}\ \text{g} \times \frac{1\ \text{kg}}{1000\ \text{g}} = \boxed{2.34 \times 10^{-7}\ \text{kg}}$$

Again, this example includes unit factors that contain only exact numbers.

✓ Problem-Solving Tip: *Conversions Within the Metric or SI System*

The SI and metric systems of units are based on powers of ten. This means that many unit conversions *within* these systems can be carried out just by shifting the decimal point. For instance, the conversion from milligrams to grams in Example 1-6 just involves shifting the decimal point to the *left* by three places. How do we know to move it to the left? We know that the gram is a larger unit of mass than the milligram, so the number of grams in a given mass must be a *smaller* number than the number of milligrams. After you carry out many such conversions using unit factors, you will probably begin to take such shortcuts. Always think about the answer, to see whether it should be larger or smaller than the quantity was before conversion.

Unity raised to *any* power is 1. *Any* unit factor raised to a power is still a unit factor, as the next example shows.

EXAMPLE 1-7 *Volume Conversion*

One liter is exactly 1000 cm^3. How many cubic inches are there in 1000 cm^3?

Plan

We would multiply by the unit factor $\frac{1 \text{ in.}}{2.54 \text{ cm}}$ to convert cm to in. Here we require the *cube* of this unit factor.

Solution

$$\underline{?}\ \text{in.}^3 = 1000\ \text{cm}^3 \times \left(\frac{1\ \text{in.}}{2.54\ \text{cm}}\right)^3 = 1000\ \text{cm}^3 \times \frac{1\ \text{in.}}{16.4\ \text{cm}^3} = \boxed{61.0\ \text{in.}^3}$$

Suppose we start with the equality

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

We can perform the same operation on both sides of the equation. Let's cube both sides:

$$(1\ \text{in.})^3 = (2.54\ \text{cm})^3 = 16.4\ \text{cm}^3$$

so the quantity

$$\frac{(1\ \text{in.})^3}{(2.54\ \text{cm})^3} = 1$$

is a unit factor.

Example 1-7 shows that a unit factor *cubed* is still a unit factor.

EXAMPLE 1-8 *Energy Conversion*

A common unit of energy is the erg. Convert 3.74×10^{-2} erg to the SI units of energy, joules, and kilojoules. One erg is exactly 1×10^{-7} joule (J).

Plan

The definition that relates ergs and joules is used to generate the needed unit factor. The second conversion uses a unit factor that is based on the definition of the prefix *kilo-*.

Solution

$$\underline{?}\ \text{J} = 3.74 \times 10^{-2}\ \text{erg} \times \frac{1 \times 10^{-7}\ \text{J}}{1\ \text{erg}} = \boxed{3.74 \times 10^{-9}\ \text{J}}$$

$$\underline{?}\ \text{kJ} = 3.74 \times 10^{-9}\ \text{J} \times \frac{1\ \text{kJ}}{1000\ \text{J}} = \boxed{3.74 \times 10^{-12}\ \text{kJ}}$$

Conversions between the English and SI (metric) systems are conveniently made by the unit factor method. Several conversion factors are listed in Table 1-7. It may be helpful to remember one each for

length	1 in. = 2.54 cm (exact)
mass and weight	1 lb = 454 g (near sea level)
volume	1 qt = 0.946 L or 1 L = 1.06 qt

EXAMPLE 1-9 *English–Metric Conversion*

Express 1.0 gallon in milliliters.

Plan

We ask _?_ mL = 1.0 gal and multiply by the appropriate factors.

We relate
(a) gallons to quarts, then
(b) quarts to liters, and then
(c) liters to milliliters.

gallons ⟶ quarts ⟶ liters ⟶ milliliters

Solution

$$\underline{?}\text{ mL} = 1.0\text{ gal} \times \frac{4\text{ qt}}{1\text{ gal}} \times \frac{1\text{ L}}{1.06\text{ qt}} \times \frac{1000\text{ mL}}{1\text{ L}} = \boxed{3.8 \times 10^3\text{ mL}}$$

You should now work Exercise 32.

The fact that all other units cancel to give the desired unit, milliliters, shows that we used the correct unit factors. The factors 4 qt/gal and 1000 mL/L contain only exact numbers. The factor 1 L/1.06 qt contains three significant figures. Because 1.0 gal contains only two, the answer contains only two significant figures.

Examples 1-1 through 1-9 show that multiplication by one or more unit factors changes the units and the number of units, but not the amount of whatever we are calculating.

1-10 PERCENTAGE

We often use percentages to describe quantitatively how a total is made up of its parts. In Table 1-3, we described the amounts of elements present in terms of the percentage of each element.

Percentages can be treated as unit factors. For any mixture containing substance A,

$$\underset{\text{(by mass)}}{\%\text{ A}} = \frac{\text{Parts A (by mass)}}{\text{100 parts mixture (by mass)}}$$

$$\text{Mass A} \longleftrightarrow \text{Mass mixture}$$

If we say that a sample is 24.4% carbon by mass, we mean that out of every 100 parts (exactly) by mass of sample, 24.4 parts by mass are carbon. This relationship can be represented by whichever of the two unit factors we find useful:

$$\frac{24.4\text{ parts carbon}}{100\text{ parts sample}} \quad \text{or} \quad \frac{100\text{ parts sample}}{24.4\text{ parts carbon}}$$

This ratio can be expressed in terms of grams of carbon for every 100 grams of sample, pounds of carbon for every 100 pounds of sample, or any other mass or weight unit. The next example illustrates the use of dimensional analysis involving percentage.

EXAMPLE 1-10 *Percentage*

U.S. pennies made since 1982 consist of 97.6% zinc and 2.4% copper. The mass of a particular penny is measured to be 1.494 grams. How many grams of zinc does this penny contain?

Plan

From the percentage information given, we may write the required unit factor

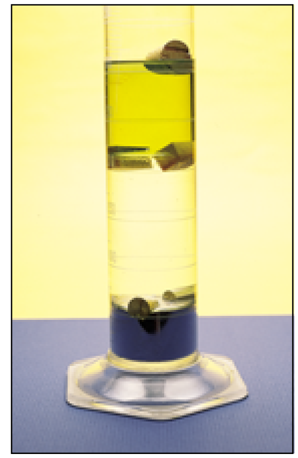
$$\frac{97.6 \text{ g zinc}}{100 \text{ g sample}}$$

Solution

$$\underline{?} \text{ g zinc} = 1.494 \text{ g sample} \times \frac{97.6 \text{ g zinc}}{100 \text{ g sample}} = \boxed{1.46 \text{ g zinc}}$$

The number of significant figures in the result is limited by the three significant figures in 97.6%. Because the definition of percentage involves *exactly* 100 parts, the number 100 is known to an infinite number of significant figures.

You should now work Exercises 59 and 60.



Six materials with different densities. The liquid layers are gasoline (*top*), water (*middle*), and mercury (*bottom*). A cork floats on gasoline. A piece of oak wood sinks in gasoline but floats on water. Brass sinks in water but floats on mercury.

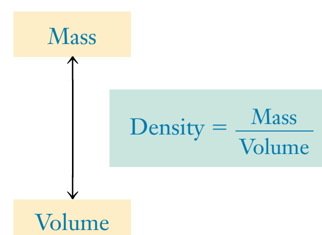
1-11 DENSITY AND SPECIFIC GRAVITY

In science, we use many terms that involve combinations of different units. Such quantities may be thought of as unit factors that can be used to convert among these units. The **density** of a sample of matter is defined as the mass per unit volume:

$$\text{density} = \frac{\text{mass}}{\text{volume}} \qquad \text{or} \qquad D = \frac{m}{V}$$

Densities may be used to distinguish between two substances or to assist in identifying a particular substance. They are usually expressed as g/cm^3 or g/mL for liquids and solids and as g/L for gases. These units can also be expressed as g·cm^{-3}, g·mL^{-1}, and g·L^{-1}, respectively. Densities of several substances are listed in Table 1-8.

See the *Saunders Interactive General Chemistry CD-ROM,* Screen 1.8, Density.



The intensive property *density* relates the two extensive properties: *mass* and *volume*.

EXAMPLE 1-11 *Density, Mass, Volume*

A 47.3-mL sample of ethyl alcohol (ethanol) has a mass of 37.32 g. What is its density?

Plan

We use the definition of density.

Solution

$$D = \frac{m}{V} = \frac{37.32 \text{ g}}{47.3 \text{ mL}} = \boxed{0.789 \text{ g/mL}}$$

You should now work Exercise 36.

TABLE 1-8 *Densities of Common Substances**

Substance	Density (g/cm^3)	Substance	Density (g/cm^3)
hydrogen (gas)	0.000089	sand*	2.32
carbon dioxide (gas)	0.0019	aluminum	2.70
cork*	0.21	iron	7.86
oak wood*	0.71	copper	8.92
ethyl alcohol	0.789	silver	10.50
water	1.00	lead	11.34
magnesium	1.74	mercury	13.59
table salt	2.16	gold	19.30

**Cork, oak wood, and sand are common materials that have been included to provide familiar reference points. They are* not *pure elements or compounds as are the other substances listed.*

These densities are given at room temperature and *one atmosphere* pressure, the average atmospheric pressure at sea level. Densities of solids and liquids change only slightly, but densities of gases change greatly, with changes in temperature and pressure.

EXAMPLE 1-12 *Density, Mass, Volume*

If 116 g of ethanol is needed for a chemical reaction, what volume of liquid would you use?

Plan

We determined the density of ethanol in Example 1-11. Here we are given the mass, m, of a sample of ethanol. So we know values for D and m in the relationship

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

We rearrange this relationship to solve for V, put in the known values, and carry out the calculation. Alternatively, we can use the unit factor method to solve the problem.

Solution

The density of ethanol is 0.789 g/mL (Table 1-8).

$$D = \frac{m}{V}, \qquad \text{so} \qquad V = \frac{m}{D} = \frac{116\text{ g}}{0.789\text{ g/mL}} = \boxed{147\text{ mL}}$$

Alternatively,

$$\underline{?}\text{ mL} = 116\text{ g} \times \frac{1\text{ mL}}{0.789\text{ g}} = \boxed{147\text{ mL}}$$

You should now work Exercise 39.

Observe that density gives two unit factors. In this case, they are $\frac{0.789\text{ g}}{1\text{ mL}}$ and $\frac{1\text{ mL}}{0.789\text{ g}}$

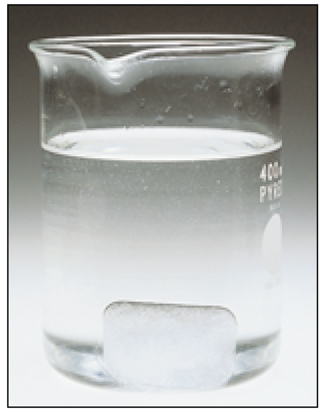
EXAMPLE 1-13 *Unit Conversion*

Express the density of mercury in lb/ft^3.

Plan

The density of mercury is 13.59 g/cm^3 (see Table 1-8). To convert this value to the desired units, we can use unit factors constructed from the conversion factors in Table 1-7.

Ice is slightly less dense than liquid water, so ice floats in water.



Solid ethyl alcohol is more dense than liquid ethyl alcohol. This is true of nearly every known substance.

Solution

$$\underline{?}\ \frac{\text{lb}}{\text{ft}^3} = 13.59\ \frac{\text{g}}{\text{cm}^3} \times \frac{1\ \text{lb}}{453.6\ \text{g}} \times \left(\frac{2.54\ \text{cm}}{1\ \text{in.}}\right)^3 \times \left(\frac{12\ \text{in.}}{1\ \text{ft}}\right)^3 = \boxed{848.4\ \text{lb/ft}^3}$$

It would take a very strong person to lift a cubic foot of mercury!

The **specific gravity** (Sp. Gr.) of a substance is the ratio of its density to the density of water, both at the same temperature.

$$\text{Sp. Gr.} = \frac{D_{\text{substance}}}{D_{\text{water}}}$$

The density of water is 1.000 g/mL at 3.98°C, the temperature at which the density of water is greatest. Variations in the density of water with changes in temperature, however, are small enough that we can use 1.00 g/mL up to 25°C without introducing significant errors into our calculations.

Density and specific gravity are both intensive properties; that is, they do not depend on the size of the sample. Specific gravities are dimensionless numbers.

EXAMPLE 1-14 *Density, Specific Gravity*

The density of table salt is 2.16 g/mL at 20°C. What is its specific gravity?

Plan

We use the preceding definition of specific gravity. The numerator and denominator have the same units, so the result is dimensionless.

Solution

$$\text{Sp. Gr.} = \frac{D_{\text{salt}}}{D_{\text{water}}} = \frac{2.16\ \text{g/mL}}{1.00\ \text{g/mL}} = \boxed{2.16}$$

This example also demonstrates that the density and specific gravity of a substance are numerically equal near room temperature if density is expressed in g/mL (g/cm^3).

Labels on commercial solutions of acids give specific gravities and the percentage by mass of the acid present in the solution. From this information, the amount of acid present in a given volume of solution can be calculated.

EXAMPLE 1-15 *Specific Gravity, Volume, Percentage by Mass*

Battery acid is 40.0% sulfuric acid, H_2SO_4, and 60.0% water by mass. Its specific gravity is 1.31. Calculate the mass of pure H_2SO_4 in 100.0 mL of battery acid.

Plan

The percentages are given on a mass basis, so we must first convert the 100.0 mL of acid solution (soln) to mass. To do this, we need a value for the density. We have demonstrated that density and specific gravity are numerically equal at 20°C because the density of water is 1.00 g/mL. We can use the density as a unit factor to convert the given volume of the solution to mass of the solution. Then we use the percentage by mass to convert the mass of the solution to the mass of the acid.

Solution

From the given value for specific gravity, we may write

$$\text{Density} = 1.31\ \text{g/mL}$$

The solution is 40.0% H_2SO_4 and 60.0% H_2O by mass. From this information we may construct the desired unit factor:

$$\frac{40.0\ \text{g}\ H_2SO_4}{100\ \text{g soln}} \longrightarrow \boxed{\text{because 100 g of solution contains 40.0 g of } H_2SO_4}$$

We can now solve the problem:

$$\underline{?}\ H_2SO_4 = 100.0\ \text{mL soln} \times \frac{1.31\ \text{g soln}}{1\ \text{mL soln}} \times \frac{40.0\ \text{g}\ H_2SO_4}{100\ \text{g soln}} = \boxed{52.4\ \text{g}\ H_2SO_4}$$

You should now work Exercise 43.

1-12 HEAT AND TEMPERATURE

See the *Saunders Interactive General Chemistry CD-ROM,* Screen 1.10, Temperature.

In Section 1-1 you learned that heat is one form of energy. You also learned that the many forms of energy can be interconverted and that in chemical processes, chemical energy is converted to heat energy or vice versa. The amount of heat a process uses (*endothermic*) or gives off (*exothermic*) can tell us a great deal about that process. For this reason it is important for us to be able to measure the intensity of heat.

Temperature measures the intensity of heat, the "hotness" or "coldness" of a body. A piece of metal at 100°C feels hot to the touch, whereas an ice cube at 0°C feels cold. Why? Because the temperature of the metal is higher, and that of the ice cube lower, than body temperature. **Heat** is a form of energy that *always flows spontaneously from a hotter body to a colder body*—never in the reverse direction.

Temperatures can be measured with mercury-in-glass thermometers. A mercury thermometer consists of a reservoir of mercury at the base of a glass tube, open to a very thin (capillary) column extending upward. Mercury expands more than most other liquids as its temperature rises. As it expands, its movement up into the evacuated column can be seen.

Anders Celsius (1701–1744), a Swedish astronomer, developed the Celsius temperature scale, formerly called the centigrade temperature scale. When we place a Celsius thermometer in a beaker of crushed ice and water, the mercury level stands at exactly 0°C, the lower reference point. In a beaker of water boiling at one atmosphere pressure, the mercury level stands at exactly 100°C, the higher reference point. There are 100 equal steps between these two mercury levels. They correspond to an interval of 100 degrees between the melting point of ice and the boiling point of water at one atmosphere. Figure 1-16 shows how temperature marks between the reference points are established.

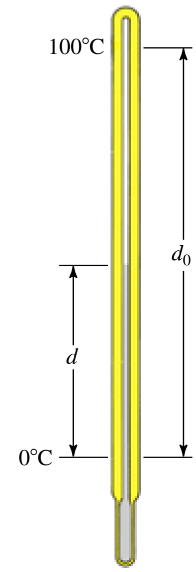


Figure 1-16 At 45°C, as read on a mercury-in-glass thermometer, d equals $0.45d_0$ where d_0 is the distance from the mercury level at 0°C to the level at 100°C.

In the United States, temperatures are frequently measured on the temperature scale devised by Gabriel Fahrenheit (1686–1736), a German instrument maker. On this scale the freezing and boiling points of water are defined as 32°F and 212°F, respectively. In scientific work, temperatures are often expressed on the **Kelvin** (absolute) temperature scale. As we shall see in Section 12-5, the zero point of the Kelvin temperature scale is *derived* from the observed behavior of all matter.

Relationships among the three temperature scales are illustrated in Figure 1-17. Between the freezing point of water and the boiling point of water, there are 100 steps (°C or kelvins, respectively) on the Celsius and Kelvin scales. Thus the "degree" is the same size on the Celsius and Kelvin scales. But every Kelvin temperature is 273.15 units above the corresponding Celsius temperature. The relationship between these two scales is as follows:

$$\underline{?}\ \text{K} = {}^\circ\text{C} + 273.15^\circ \qquad \text{or} \qquad \underline{?}\ {}^\circ\text{C} = \text{K} - 273.15^\circ$$

We shall usually round 273.15 to 273.

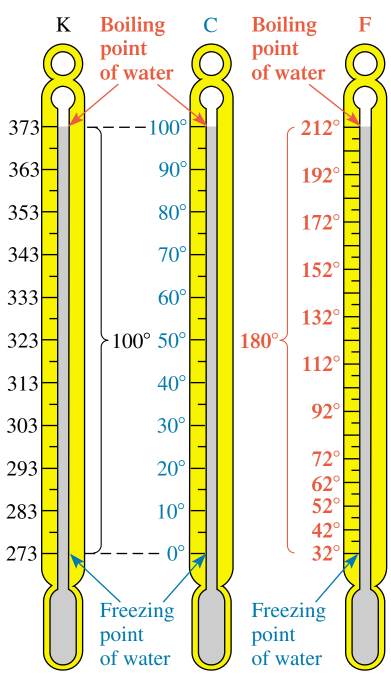


Figure 1-17 The relationships among the Kelvin, Celsius (centigrade), and Fahrenheit temperature scales.

In the SI system, "degrees Kelvin" are abbreviated simply as K rather than °K and are called **kelvins.**

Any temperature *change* has the same numerical value whether expressed on the Celsius scale or on the Kelvin scale. For example, a change from 25°C to 59°C represents a *change* of 34 Celsius degrees. Converting these to the Kelvin scale, the same change is expressed as (273 + 25) = 298 K to (59 + 273) = 332 K, or a *change* of 34 kelvins.

Comparing the Fahrenheit and Celsius scales, we find that the intervals between the same reference points are 180 Fahrenheit degrees and 100 Celsius degrees, respectively. Thus a Fahrenheit degree must be smaller than a Celsius degree. It takes 180 Fahrenheit degrees to cover the same temperature *interval* as 100 Celsius degrees. From this information, we can construct the unit factors for temperature *changes*:

The numbers in these ratios are *exact* numbers, so they do not affect the number of significant figures in the calculated result.

$$\frac{180°F}{100°C} \quad \text{or} \quad \frac{1.8°F}{1.0°C} \quad \text{and} \quad \frac{100°C}{180°F} \quad \text{or} \quad \frac{1.0°C}{1.8°F}$$

But the starting points of the two scales are different, so we *cannot convert* a temperature on one scale to a temperature on the other just by multiplying by the unit factor. In converting from °F to °C, we must subtract 32 Fahrenheit degrees to reach the zero point on the Celsius scale (Figure 1-17).

$$\underline{?}\ °F = \left(x°C \times \frac{1.8°F}{1.0°C}\right) + 32°F \qquad \text{and} \qquad \underline{?}\ °C = \frac{1.0°C}{1.8°F}(x°F - 32°F)$$

These are often remembered in abbreviated form:

$$°F = 1.8°C + 32°$$

$$°C = \frac{(°F - 32°)}{1.8}$$

Either of these equations can be rearranged to obtain the other one, so you need to learn only one of them.

EXAMPLE 1-16 *Temperature Conversion*

When the temperature reaches "100.°F in the shade," it's hot. What is this temperature on the Celsius scale?

Plan

We use the relationship $\underline{?}\ °C = \frac{1.0°C}{1.8°F}(x°F - 32°F)$ to carry out the desired conversion.

Solution

A temperature of 100.°F is 38°C.

$$\underline{?}\ °C = \frac{1.0°C}{1.8°F}(100.°F - 32°F) = \frac{1.0°C}{1.8°F}(68°F) = \boxed{38°C}$$

EXAMPLE 1-17 *Temperature Conversion*

When the absolute temperature is 400 K, what is the Fahrenheit temperature?

Plan

We first use the relationship $\underline{?}\ °C = K - 273°$ to convert from kelvins to degrees Celsius, then we carry out the further conversion from degrees Celsius to degrees Fahrenheit.

Solution

$$\underline{?}\ °C = (400\ K - 273\ K)\frac{1.0°C}{1.0\ K} = 127°C$$

$$\underline{?}\ °F = \left(127°C \times \frac{1.8°F}{1.0°C}\right) + 32°F = \boxed{261°F}$$

You should now work Exercise 46.

1-13 HEAT TRANSFER AND THE MEASUREMENT OF HEAT

Chemical reactions and physical changes occur with either the simultaneous evolution of heat **(exothermic processes)** or the absorption of heat **(endothermic processes).** The amount of heat transferred in a process is usually expressed in joules or in calories.

The SI unit of energy and work is the **joule (J),** which is defined as 1 $kg \cdot m^2/s^2$. The kinetic energy (KE) of a body of mass m moving at speed v is given by $\frac{1}{2}mv^2$. A 2-kg object moving at one meter per second has KE $= \frac{1}{2}(2\ kg)(1\ m/s)^2 = 1\ kg \cdot m^2/s^2 = 1$ J. You may find it more convenient to think in terms of the amount of heat required to raise the temperature of one gram of water from 14.5°C to 15.5°C, which is 4.184 J.

In English units this corresponds to a 4.4-pound object moving at 197 feet per minute, or 2.2 miles per hour. In terms of electrical energy, one joule is equal to one watt · second. Thus, one joule is enough energy to operate a 10-watt light bulb for $\frac{1}{10}$ second.

One **calorie** is defined as exactly 4.184 J. The so-called "large calorie," used to indicate the energy content of foods, is really one kilocalorie, that is, 1000 calories. We shall do most calculations in joules.

The calorie was originally defined as the amount of heat necessary to raise the temperature of one gram of water at one atmosphere from 14.5°C to 15.5°C.

The **specific heat** of a substance is the amount of heat required to raise the temperature of one gram of the substance one degree Celsius (also one kelvin) with no change in phase. Changes in phase (physical state) absorb or liberate relatively large amounts of energy (see Figure 1-5). The specific heat of each substance, a physical property, is different for the solid, liquid, and gaseous phases of the substance. For example, the specific heat of ice is 2.09 J/g·°C near 0°C; for liquid water it is 4.18 J/g·°C; and for steam it is 2.03 J/g·°C near 100°C. The specific heat for water is quite high. A table of specific heats is provided in Appendix E.

$$\text{Specific heat} = \frac{(\text{amount of heat in J})}{(\text{mass of substance in g})(\text{temperature change in °C})}$$

The specific heat of a substance varies *slightly* with temperature and pressure. These variations can be ignored for calculations in this text.

The units of specific heat are $\frac{J}{g \cdot °C}$ or $J \cdot g^{-1} \cdot °C^{-1}$.

The **heat capacity** of a body is the amount of heat required to raise its temperature 1°C. The heat capacity of a body is its mass in grams times its specific heat. The heat capacity refers to the mass of that particular body, so its units do not include mass. The units are J/°C or $J \cdot °C^{-1}$.

EXAMPLE 1-18 *Specific Heat*

How much heat, in joules, is required to raise the temperature of 205 g of water from 21.2°C to 91.4°C?

In this example, we calculate the amount of heat needed to prepare a cup of hot tea.

Plan

The specific heat of a substance is the amount of heat required to raise the temperature of 1 g of substance 1°C:

$$\text{Specific heat} = \frac{(\text{amount of heat in J})}{(\text{mass of substance in g})(\text{temperature change in °C})}$$

We can rearrange the equation so that

$$(\text{Amount of heat}) = (\text{mass of substance})\ (\text{specific heat})\ (\text{temperature change})$$

Alternatively, we can use the unit factor approach.

Solution

$$\text{Amount of heat} = (205\ g)\ (4.18\ J/g \cdot °C)\ (70.2°C) = \boxed{6.02 \times 10^4\ J}$$

By the unit factor approach,

$$\text{Amount of heat} = (205 \text{ g})\left(\frac{4.18 \text{ J}}{1 \text{ g}\cdot°\text{C}}\right)(70.2°\text{C}) = \boxed{6.02 \times 10^4 \text{ J}} \quad \text{or} \quad \boxed{60.2 \text{ kJ}}$$

All units except joules cancel. To cool 205 g of water from 91.4°C to 21.2°C, it would be necessary to remove exactly the same amount of heat, 60.2 kJ.

You should now work Exercises 54 and 55.

When two objects at different temperatures are brought into contact, heat flows from the hotter to the colder body (Figure 1-18); this continues until the two are at the same temperature. We say that the two objects are then in *thermal equilibrium.* The temperature change that occurs for each object depends on the initial temperatures and the relative masses and specific heats of the two materials.

EXAMPLE 1-19 *Specific Heat*

A 385-gram chunk of iron is heated to 97.5°C. Then it is immersed in 247 grams of water originally at 20.7°C. When thermal equilibrium has been reached, the water and iron are both at 31.6°C. Calculate the specific heat of iron.

Plan

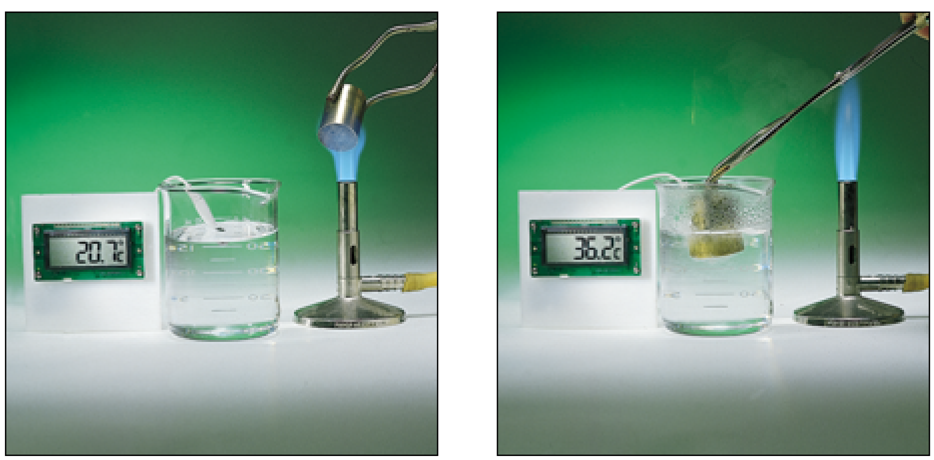
The amount of heat gained by the water as it is warmed from 20.7°C to 31.6°C is the same as the amount of heat lost by the iron as it cools from 97.5°C to 31.6°C. We can equate these two amounts of heat and solve for the unknown specific heat.

Solution

In specific heat calculations, we use the *magnitude* of the temperature change (i.e., a positive number), so we subtract the lower temperature from the higher one in both cases.

$$\text{Temperature change of water} = 31.6°\text{C} - 20.7°\text{C} = 10.9°\text{C}$$

$$\text{Temperature change of iron} = 97.5°\text{C} - 31.6°\text{C} = 65.9°\text{C}$$



(a) **(b)**

Figure 1-18 A hot object, such as a heated piece of metal (a), is placed into cooler water. Heat is transferred from the hotter metal bar to the cooler water until the two reach the same temperature (b). We say that they are then at *thermal equilibrium.*

$$\text{Number of joules gained by water} = (247\text{ g})\left(4.18\,\frac{\text{J}}{\text{g}\cdot°\text{C}}\right)(10.9°\text{C})$$

Let x = specific heat of iron

$$\text{Number of joules lost by iron} = (385\text{ g})\left(x\,\frac{\text{J}}{\text{g}\cdot°\text{C}}\right)(65.9°\text{C})$$

We set these two quantities equal to one another and solve for x.

$$(247\text{g})\left(4.18\,\frac{\text{J}}{\text{g}\cdot°\text{C}}\right)(10.9°\text{C}) = (385\text{ g})\left(x\,\frac{\text{J}}{\text{g}\cdot°\text{C}}\right)(65.9°\text{C})$$

$$x = \frac{(247\text{ g})\left(4.18\,\frac{\text{J}}{\text{g}\cdot°\text{C}}\right)(10.9°\text{C})}{(385\text{ g})(65.9°\text{C})} = \boxed{0.444\,\frac{\text{J}}{\text{g}\cdot°\text{C}}}$$

You should now work Exercise 58.

The specific heat of iron is much smaller than the specific heat of water.

$$\frac{\text{Specific heat of iron}}{\text{Specific heat of water}} = \frac{0.444\text{ J/g}\cdot°\text{C}}{4.18\text{ J/g}\cdot°\text{C}} = 0.106$$

The amount of heat required to raise the temperature of 205 g of iron by 70.2°C (as we calculated for water in Example 1-18) is

$$\text{Amount of heat} = (205\text{ g})\left(\frac{0.444\text{ J}}{\text{g}\cdot°\text{C}}\right)(70.2°\text{C}) = 6.39 \times 10^3\text{ J, or }6.39\text{ kJ}$$

We see that the amount of heat required to accomplish a given change in temperature for a given quantity of iron is less than that for the same quantity of water, by the same ratio.

$$\frac{\text{Number of joules required to warm 205 g of iron by 70.2°C}}{\text{Number of joules required to warm 205 g of water by 70.2°C}} = \frac{6.39\text{ kJ}}{60.2\text{ kJ}} = 0.106$$

It might not be necessary to carry out explicit calculations when we are looking only for qualitative comparisons.

EXAMPLE 1-20 *Comparing Specific Heats*

We add the same amount of heat to 10.0 grams of each of the following substances starting at 20.0°C: liquid water, $H_2O(\ell)$; liquid mercury; $Hg(\ell)$; liquid benzene, $C_6H_6(\ell)$; and solid aluminum, Al(s). Rank the samples from lowest to highest final temperature. Refer to Appendix E for required data.

Plan

We can obtain the values of specific heats (Sp. Ht.) for these substances from Appendix E. The higher the specific heat for a substance, the more heat is required to raise a given mass of sample by a given temperature change, so the less its temperature changes by a given amount of heat. The substance with the lowest specific heat undergoes the largest temperature change, and the one with the highest specific heat undergoes the smallest temperature change. *It is not necessary to calculate the amount of heat required to answer this question.*

Solution

The specific heats obtained from Appendix E are as follows:

Substance	Sp. Ht. $\left(\frac{J}{g \cdot °C}\right)$
$H_2O(\ell)$	4.18
$Hg(\ell)$	0.138
$C_6H_6(\ell)$	1.74
$Al(s)$	0.900

Ranked from highest to lowest specific heats: $H_2O(\ell) > C_6H_6(\ell) > Al(s) > Hg(\ell)$. Adding the same amount of heat to the same size sample of these substances changes the temperature of $H_2O(\ell)$ the least and that of $Hg(\ell)$ the most. The ranking from lowest to highest final temperature is

$$H_2O(\ell) < C_6H_6(\ell) < Al(s) < Hg(\ell)$$

You should now work Exercise 70.

Key Terms

Accuracy How closely a measured value agrees with the correct value.

Calorie Defined as exactly 4.184 joules. Originally defined as the amount of heat required to raise the temperature of one gram of water from 14.5°C to 15.5°C.

Chemical change A change in which one or more new substances are formed.

Chemical property See *Properties.*

Compound A substance composed of two or more elements in fixed proportions. Compounds can be decomposed into their constituent elements.

Density Mass per unit volume, $D = m/V$.

Element A substance that cannot be decomposed into simpler substances by chemical means.

Endothermic Describes processes that absorb heat energy.

Energy The capacity to do work or transfer heat.

Exothermic Describes processes that release heat energy.

Extensive property A property that depends on the amount of material in a sample.

Heat A form of energy that flows between two samples of matter because of their difference in temperature.

Heat capacity The amount of heat required to raise the temperature of a body (of whatever mass) one degree Celsius.

Heterogeneous mixture A mixture that does not have uniform composition and properties throughout.

Homogeneous mixture A mixture that has uniform composition and properties throughout.

Intensive property A property that is independent of the amount of material in a sample.

Joule A unit of energy in the SI system. One joule is 1 $kg \cdot m^2/s^2$, which is also 0.2390 cal.

Kinetic energy Energy that matter possesses by virtue of its motion.

Law of Conservation of Energy Energy cannot be created or destroyed in a chemical reaction or in a physical change; it may be changed from one form to another.

Law of Conservation of Matter No detectable change occurs in the total quantity of matter during a chemical reaction or during a physical change.

Law of Conservation of Matter and Energy The combined amount of matter and energy available in the universe is fixed.

Law of Constant Composition See *Law of Definite Proportions.*

Law of Definite Proportions Different samples of any pure compound contain the same elements in the same proportions by mass; also known as the *Law of Constant Composition.*

Mass A measure of the amount of matter in an object. Mass is usually measured in grams or kilograms.

Matter Anything that has mass and occupies space.

Mixture A sample of matter composed of variable amounts of two or more substances, each of which retains its identity and properties.

Physical change A change in which a substance changes from one physical state to another, but no substances with different compositions are formed.

Physical property See *Properties.*

Potential energy Energy that matter possesses by virtue of its position, condition, or composition.

Precision How closely repeated measurements of the same quantity agree with one another.

Properties Characteristics that describe samples of matter. Chemical properties are exhibited as matter undergoes chemical changes. Physical properties are exhibited by matter with no changes in chemical composition.

Scientific (natural) law A general statement based on the observed behavior of matter, to which no exceptions are known.

Significant figures Digits that indicate the precision of measurements—digits of a measured number that have uncertainty only in the last digit.

Specific gravity The ratio of the density of a substance to the density of water at the same temperature.

Specific heat The amount of heat required to raise the temperature of one gram of a substance one degree Celsius.

Substance Any kind of matter all specimens of which have the same chemical composition and physical properties.

Symbol (of an element) A letter or group of letters that represents (identifies) an element.

Temperature A measure of the intensity of heat, that is, the hotness or coldness of a sample or object.

Unit factor A factor in which the numerator and denominator are expressed in different units but represent the same or equivalent amounts. Multiplying by a unit factor is the same as multiplying by one.

Weight A measure of the gravitational attraction of the earth for a body.

Exercises

Asterisks are used to denote some of the more challenging exercises.

Matter and Energy

1. Define the following terms, and illustrate each with a specific example: (a) matter; (b) energy; (c) mass; (d) exothermic process; (e) intensive property.

2. Define the following terms, and illustrate each with a specific example: (a) weight; (b) potential energy; (c) kinetic energy; (d) endothermic process; (e) extensive property.

3. State the following laws, and illustrate each.
(a) the Law of Conservation of Matter
(b) the Law of Conservation of Energy
(c) the Law of Conservation of Matter and Energy

4. List the three states of matter and some characteristics of each. How are they alike? different?

5. An incandescent light bulb functions because of the flow of electric current. Does the incandescent light bulb convert all of the electrical energy to light? Observe a functioning incandescent light bulb, and explain what occurs with reference to the Law of Conservation of Energy.

6. All electrical motors are less than 100% efficient in converting electrical energy into useable work. How can their efficiency be less than 100% and the Law of Conservation of Energy still be valid?

States of Matter

7. What is a homogeneous mixture? Which of the following are pure substances? Which of the following are homogeneous mixtures? Explain your answers. (a) sugar dissolved in water; (b) tea and ice; (c) french onion soup; (d) mud; (e) gasoline; (f) carbon dioxide; (g) a chocolate-chip cookie.

8. Define the following terms clearly and concisely. Give two illustrations of each: (a) substance; (b) mixture; (c) element; (d) compound.

9. Classify each of the following as an element, a compound, or a mixture. Justify your classification: (a) a soft drink; (b) water; (c) air; (d) chicken noodle soup; (e) table salt; (f) popcorn; (g) aluminum foil.

10. Classify each of the following as an element, a compound, or a mixture. Justify your classification: (a) coffee; (b) silver; (c) calcium carbonate; (d) ink from a ballpoint pen; (e) toothpaste.

11. Sand, candle wax, and table sugar are placed in a beaker and stirred. (a) Is the resulting combination a mixture? If so, what kind of mixture? (b) Design an experiment in which the sand, candle wax, and table sugar can be separated.

12. A $10 gold piece minted in the early part of the 1900s appeared to have a dirty area. The dirty appearance could not be removed by careful cleaning. Close examination of the coin revealed that the "dirty" area was really pure copper. Is the mixture of gold and copper in this coin a heterogeneous or homogeneous mixture?

Chemical and Physical Properties

13. Distinguish between the following pairs of terms and give two specific examples of each: (a) chemical properties and physical properties; (b) intensive properties and extensive properties; (c) chemical changes and physical changes; (d) mass and weight.

14. Which of the following are chemical properties, and which are physical properties? (a) Baking powder gives off bubbles of carbon dioxide when added to water. (b) A particular type of steel consists of 95% iron, 4% carbon, and 1% miscellaneous other elements. (c) The density of gold is 19.3 g/mL. (d) Iron dissolves in hydrochloric acid with the evolution of hydrogen gas. (e) Fine steel wool burns in air. (f) Refrigeration slows the rate at which fruit ripens.

15. Which of the following are chemical properties, and which are physical properties? (a) Metallic sodium is soft enough to be cut with a knife. (b) When sodium metal is cut, the surface is at first shiny; after a few seconds of exposure to air, it turns a dull gray. (c) The density of sodium is 0.97 g/mL. (d) Cork floats on water. (e) When sodium comes in contact with water, it melts, evolves a flammable gas, and eventually disappears altogether. (f) Household bleach changes the color of your favorite T-shirt from purple to pink.

16. Describe each of the following as a chemical change, a physical change, or both. (a) A wet towel dries in the sun. (b) Lemon juice is added to tea, causing its color to change. (c) Hot air rises over a radiator. (d) Coffee is brewed by passing hot water through ground coffee. (e) Dynamite explodes.

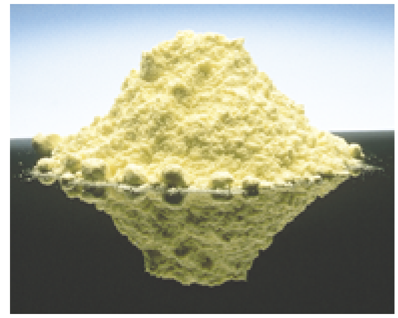
17. Describe each of the following as a chemical change, a physical change, or both. (a) Powdered sulfur is heated, first melting and then burning. (b) Alcohol is evaporated by heating. (c) Transparent rock candy (pure sugar crystals) is finely ground into an opaque white powder. (d) Chlorine gas is bubbled through concentrated sea water, releasing liquid bromine. (e) Electricity is passed through water, resulting in the evolution of hydrogen and oxygen gases. (f) An ice cube in your glass of water melts.

18. Which of the following processes are exothermic? endothermic? How can you tell? (a) combustion; (b) freezing water; (c) melting ice; (d) boiling water; (e) condensing steam; (f) burning paper.

19. Which of the following illustrate the concept of potential energy and which illustrate kinetic energy? (a) a spinning gyroscope; (b) a rubber band stretched around a newspaper; (c) a frozen pint of ice cream; (d) a comet moving through space; (e) a basketball dropping through a net; (f) the roof of a house.

20. Which of the following illustrate the concept of potential energy and which illustrate kinetic energy? (a) the gasoline in a car's gas tank; (b) a car's battery; (c) the car as it moves along the highway; (d) a space vehicle in orbit about the earth; (e) a river flowing; (f) fatty tissue in your body.

21. A sample of yellow sulfur powder is placed in a sealed flask with the air removed and replaced with an inert gas. Heat is applied by means of a flame from a Bunsen burner until the sulfur melts and begins to boil. After cooling, the material in the flask is reddish and has the consistency of used chewing gum. Careful chemical analysis tells us that the substance is pure sulfur. Is this a chemical or physical change? Propose an explanation for the change.



22. A weighed sample of yellow sulfur is placed in a flask. The flask is gently heated using a Bunsen burner. Observation indicates that nothing appears to happen to the sulfur during the heating, but the mass of sulfur is less than before the heating and there is a sharp odor that was not present before the heating. Propose an explanation of what caused the change in the mass of the sulfur. Is your hypothesis of the mass change a chemical or physical change?

Measurements and Calculations

23. Express the following numbers in scientific notation: (a) 6500.; (b) 0.00630; (c) 860 (assume that this number is measured to ± 10); (d) 860 (assume that this number is measured to ± 1); (e) 186,000; (f) 0.10010.

24. Express the following exponentials as ordinary numbers: (a) 5.26×10^4; (b) 4.10×10^{-6}; (c) 16.00×10^2; (d) 8.206×10^{-4}; (e) 9.346×10^3; (f) 9.346×10^{-3}.

25. Which of the following are likely to be exact numbers? Why? (a) 554 in.; (b) 7 computers; (c) $20,355.47; (d) 25 lb of sugar; (e) 12.5 gal of diesel fuel; (f) 5446 ants.

26. To which of the quantities appearing in the following statements would the concept of significant figures apply? Where it would apply, indicate the number of significant figures. (a) The density of platinum at 20°C is 21.45 g/cm^3. (b) Wilbur Shaw won the Indianapolis 500-mile race in 1940 with an average speed of 114.277 mi/h. (c) A mile is defined as 5280 ft. (d) The International Committee for Weights and Measures "accepts that the curie be . . . retained as a unit of radioactivity, with the value $3.7 \times 10^{10}\ s^{-1}$." (This resolution was passed in 1964.)

27. The circumference of a circle is given by πd, where d is the diameter of the circle. Calculate the circumference of a circle with a diameter of 6.91 cm. Use the value of 3.141593 for π.

28. What is the total weight of 75 cars weighing an average of 1532.5 lb?

29. Indicate the multiple or fraction of 10 by which a quantity is multiplied when it is preceded by each of the following prefixes. (a) M; (b) m; (c) c; (d) d; (e) k; (f) n.

30. Carry out each of the following conversions. (a) 18.5 m to km; (b) 16.3 km to m; (c) 247 kg to g; (d) 4.32 L to mL; (e) 85.9 dL to L; (f) 8251 L to cm^3.

31. Express 5.31 centimeters in meters, millimeters, kilometers, and micrometers.

32. Express (a) 1.00 ft^3 in liters; (b) 1.00 L in pints; (c) miles per gallon in kilometers per liter.

33. The screen of a laptop computer measures 8.25 in. wide and 6.25 in. tall. If this computer were being sold in Europe, what would be the metric size of the screen used in the specifications for the computer?

34. If the price of gasoline is \$1.229/gal, what is its price in cents per liter?

35. Suppose your automobile gas tank holds 16 gal and the price of gasoline is \$0.325/L. How much would it cost to fill your gas tank?

36. What is the density of silicon, if 50.6 g occupies 21.72 mL?

37. What is the mass of a rectangular piece of copper 24.4 cm $\times$ 11.4 cm $\times$ 7.9 cm? The density of copper is 8.92 g/cm^3.

38. A small crystal of sucrose (table sugar) had a mass of 5.536 mg. The dimensions of the box-like crystal were 2.20 mm $\times$ 1.36 mm $\times$ 1.12 mm. What is the density of sucrose expressed in g/cm^3?

39. Vinegar has a density of 1.0056 g/cm^3. What is the mass of three L of vinegar?

40. The density of silver is 10.5 g/cm^3. (a) What is the volume, in cm^3, of an ingot of silver with mass 0.615 kg? (b) If this sample of silver is a cube, how long is each edge in cm? (c) How long is the edge of this cube in inches?

***41.** A container has a mass of 73.91 g when empty and 91.44 g when filled with water. The density of water is 1.0000 g/cm^3. (a) Calculate the volume of the container. (b) When filled with an unknown liquid, the container had a mass of 88.42 g. Calculate the density of the unknown liquid.

***42.** The mass of an empty container is 77.664 g. The mass of the container filled with water is 99.646 g. (a) Calculate the volume of the container, using a density of 1.0000 g/cm^3 for water. (b) A piece of metal was added to the empty container, and the combined mass was 85.308 g. Calculate the mass of the metal. (c) The container with the metal was filled with water, and the mass of the entire system was 106.442 g. What mass of water was added? (d) What volume of water was added? (e) What is the volume of the piece of metal? (f) Calculate the density of the metal.

43. A solution is 40.0% acetic acid (the characteristic component in vinegar) by mass. The density of this solution is 1.049 g/mL at 20°C. Calculate the mass of pure acetic acid in 100.0 mL of this solution at 20°C.

Heat Transfer and Temperature Measurement

44. Which represents a larger temperature interval: (a) a Celsius degree or a Fahrenheit degree? (b) a kelvin or a Fahrenheit degree?

45. Express (a) 283°C in K; (b) 15.25 K in °C; (c) -32.0°C in °F; (d) 100.0°F in K.

46. Express (a) 0°F in °C; (b) 98.6°F in K; (c) 298 K in °F; (d) 11.3°C in °F.

47. Make each of the following temperature conversions: (a) 27°C to °F, (b) -27°C to °F, and (c) 100°F to °C.

***48.** On the Réamur scale, which is no longer used, water freezes at 0°R and boils at 80°R. (a) Derive an equation that relates this to the Celsius scale. (b) Derive an equation that relates this to the Fahrenheit scale. (c) Mercury is a liquid metal at room temperature. It boils at 356.6°C (673.9°F). What is the boiling point of mercury on the Réamur scale?

49. Liquefied gases have boiling points well below room temperature. On the Kelvin scale the boiling points of the following gases are: He, 4.2 K; N_2, 77.4 K. Convert these temperatures to the Celsius and the Fahrenheit scales.

50. Convert the temperatures at which the following metals melt to the Celsius and Fahrenheit scales: Al, 933.6 K; Ag, 1235.1 K.

51. What is the melting point of lead in °F (mp = 327.5°C)?

52. The average temperature of a healthy German shepherd is 101.5°F. Express this temperature in degrees Celsius. Express this temperature in kelvins.

53. Calculate the amount of heat required to raise the temperature of 78.2 g of water from 10.0°C to 35.0°C. The specific heat of water is 4.18 J/g·°C.

54. The specific heat of aluminum is 0.895 J/g·°C. Calculate the amount of heat required to raise the temperature of 22.1 g of aluminum from 27.0°C to 65.5°C.

55. How much heat must be removed from 15.5 g of water at 90.0°C to cool it to 43.2°C?

***56.** In some solar-heated homes, heat from the sun is stored in rocks during the day and then released during the cooler night. (a) Calculate the amount of heat required to raise the temperature of 78.7 kg of rocks from 25.0°C to 43.0°C. Assume that the rocks are limestone, which is essentially pure calcium carbonate. The specific heat of calcium carbonate is 0.818 J/g·°C. (b) Suppose that when the rocks in part (a) cool to 30.0°C, all the heat released goes to warm the 10,000 ft^3 (2.83 $\times$ 10^5 L) of air in the house, originally at 10.0°C. To what final temperature would the air be heated? The specific heat of air is 1.004 J/g·°C, and its density is 1.20 $\times$ 10^{-3} g/mL.

***57.** A small immersion heater is used to heat water for a cup of coffee. We wish to use it to heat 235 mL of water (about a teacupful) from 25°C to 90°C in 2.00 min. What must be the heat rating of the heater, in kJ/min, to accomplish this? Ignore the heat that goes to heat the cup itself. The density of water is 0.997 g/mL.

58. When 75.0 grams of metal at 75.0°C is added to 150 grams of water at 15.0°C, the temperature of the water rises to 18.3°C. Assume that no heat is lost to the surroundings. What is the specific heat of the metal?

Mixed Exercises

59. A sample is marked as containing 22.8% calcium carbonate by mass. (a) How many grams of calcium carbonate are contained in 64.33 g of the sample? (b) How many grams of the sample would contain 11.4 g of calcium carbonate?

60. An iron ore is found to contain 9.24% hematite (a compound that contains iron). (a) How many tons of this ore would contain 8.40 tons of hematite? (b) How many kilograms of this ore would contain 9.40 kg of hematite?

***61.** A foundry releases 6.0 tons of gas into the atmosphere each day. The gas contains 2.2% sulfur dioxide by mass. What mass of sulfur dioxide is released in a five-day week?

***62.** A certain chemical process requires 215 gal of pure water each day. The available water contains 11 parts per million (ppm) by mass of salt (i.e., for every 1,000,000 parts of available water, 11 parts of salt). What mass of salt must be removed each day? A gallon of water weighs 3.67 kg.

***63.** The radius of a hydrogen atom is about 0.37 Å, and the average radius of the earth's orbit around the sun is about 1.5×10^8 km. Find the ratio of the average radius of the earth's orbit to the radius of the hydrogen atom.

64. A notice on a bridge informs drivers that the height of the bridge is 26.5 ft. How tall in meters is an 18-wheel tractor-trailer combination if it just touches the bridge?

65. Some American car manufacturers install speedometers that indicate speed in the English system and in the metric system (mi/h and km/h). What is the metric speed if the car is traveling at 65 mi/h?

***66.** The lethal dose of potassium cyanide (KCN) taken orally is 1.6 mg/kg of body weight. Calculate the lethal dose of potassium cyanide taken orally by a 175-lb person.

67. Suppose you ran a mile in 4.00 min. (a) What would be your average speed in km/h? (b) What would be your average speed in cm/s? (c) What would be your time (in minutes:seconds) for 1500 m?

68. The distance light travels through space in one year is called one light-year. Using the speed of light in a vacuum (listed in Appendix D), and assuming that one year is 365 days, determine the distance of a light-year in kilometers and in miles.

CONCEPTUAL EXERCISES

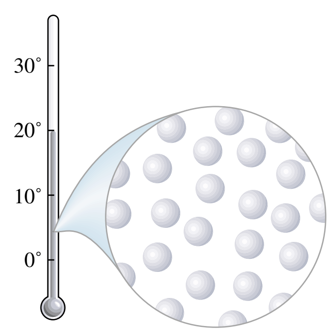
69. If you were given the job of choosing the materials from which pots and pans were to be made, what kinds of materials would you choose on the basis of specific heat? Why?

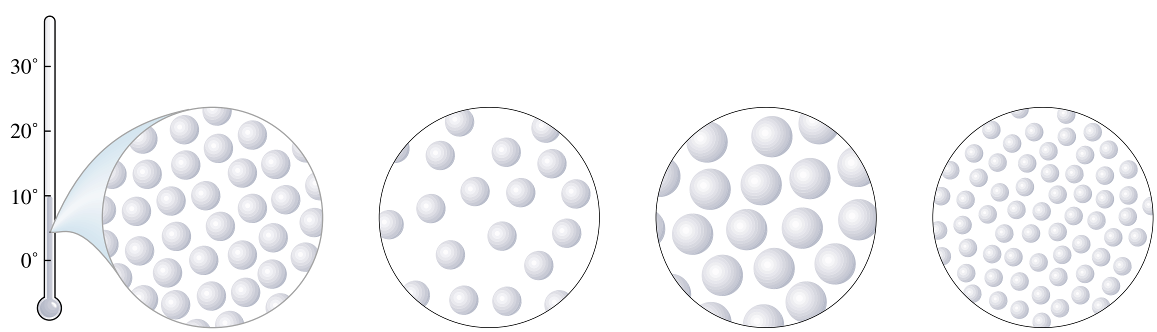
70. Fill a kitchen pan half-full with water. Place the pan on an active burner. Place one finger on the edge of the pan and one just in the water. Which finger will feel the heat first? Explain your answer.

71. Which is more dense at 0°C, ice or water? How do you know?

72. Which has the higher temperature, a sample of water at 65°C or a sample of iron at 65°F?

73. The drawing in the circle (below) is a greatly expanded representation of the molecules in the liquid of the thermometer on the left. The thermometer registers 20°C. Which of the figures (a–d) is the best representation of the liquid in this same thermometer at 10°C? (Assume that the same volume of liquid is shown in each expanded representation.)





(a) (b) (c) (d)

Exercise 73

74. During the past several years, you have gained chemical vocabulary and understanding from a variety of academic and entertainment venues. List three events that occurred early in the development of your current chemical knowledge.

***75.** At what temperature will a Fahrenheit thermometer give: (a) the same reading as a Celsius thermometer? (b) a reading that is twice that on the Celsius thermometer? (c) a reading that is numerically the same but opposite in sign from that on the Celsius thermometer?

***76.** Cesium atoms are the largest naturally occurring atoms. The radius of a cesium atom is 2.62 Å. How many cesium atoms would have to be laid side by side to give a row of cesium atoms 1.00 in. long? Assume that the atoms are spherical.

77. Based on what you have learned during the study of this chapter, write a question that requires you to use chemical information but no mathematical calculations.

78. As you write out the answer to an end-of-chapter exercise, what chemical changes occur? Did your answer involve knowledge not covered in Chapter 1?

79. *Combustion* is discussed later in this textbook; however, you probably already know what the term means. (Look it up to be sure.) List two other chemical terms that were in your vocabulary before you read Chapter 1.