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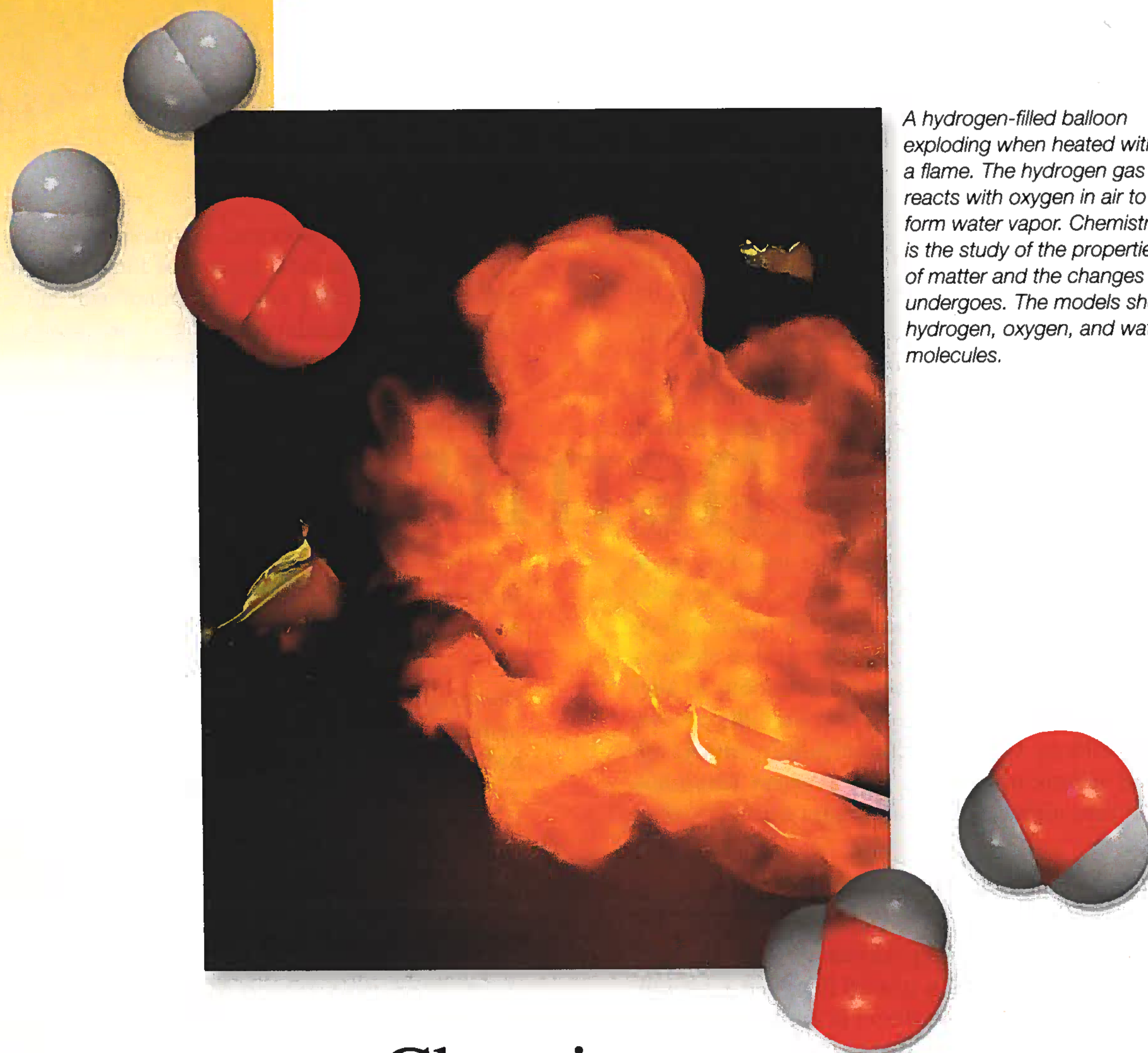


A hydrogen-filled balloon exploding when heated with a flame. The hydrogen gas reacts with oxygen in air to form water vapor. Chemistry is the study of the properties of matter and the changes it undergoes. The models show hydrogen, oxygen, and water molecules.

Chemistry

The Study of Change

A LOOK AHEAD

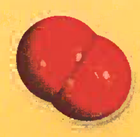
- We begin with a brief introduction to the study of chemistry and describe its role in our modern society. (1.1 and 1.2)
- Next, we become familiar with the scientific method, which is a systematic approach to research in all scientific disciplines. (1.3)
- We define matter and note that a pure substance can either be an element or a compound. We distinguish between a homogeneous mixture and a heterogeneous mixture. We also learn that, in principle, all matter can exist in one of three states: solid, liquid, and gas. (1.4 and 1.5)
- To characterize a substance, we need to know its physical properties, which can be observed without changing its identity and chemical properties, which can be demonstrated only by chemical changes. (1.6)
- Being an experimental science, chemistry involves measurements. We learn the basic SI units and use the SI-derived units for quantities like volume and density. We also become familiar with the three temperature scales: Celsius, Fahrenheit, and Kelvin. (1.7)
- Chemical calculations often involve very large or very small numbers and a convenient way to deal with these numbers is the scientific notation. In calculations or measurements, every quantity must show the proper number of significant figures, which are the meaningful digits. (1.8)
- Finally, we learn that dimensional analysis is useful in chemical calculations. By carrying the units through the entire sequence of calculations, all the units will cancel except the desired one. (1.9)

Interactive Activity Summary

1. Interactivity: Substances and Mixtures (1.4)
2. Interactivity: Elements (1.4)
3. Interactivity: SI Base Units (1.7)
4. Interactivity: Unit Prefixes (1.7)
5. Interactivity: Density (1.7)
6. Interactivity: Accuracy and Precision (1.8)
7. Interactivity: Dimensional Analysis (1.9)

Chemistry is an active, evolving science that has vital importance to our world, in both the realm of nature and the realm of society. Its roots are ancient, but as we will soon see, chemistry is every bit a modern science.

We will begin our study of chemistry at the macroscopic level, where we can see and measure the materials of which our world is made. In this chapter we will discuss the scientific method, which provides the framework for research not only in chemistry but in all other sciences as well. Next we will discover how scientists define and characterize matter. Then we will familiarize ourselves with the systems of measurement used in the laboratory. Finally, we will spend some time learning how to handle numerical results of chemical measurements and solve numerical problems. In Chapter 2 we will begin to explore the microscopic world of atoms and molecules.



化学

The Chinese characters for chemistry mean "The study of change."

1.1 Chemistry: A Science for the Twenty-First Century

Chemistry is *the study of matter and the changes it undergoes.* Chemistry is often called the central science, because a basic knowledge of chemistry is essential for students of biology, physics, geology, ecology, and many other subjects. Indeed, it is central to our way of life; without it, we would be living shorter lives in what we would consider primitive conditions, without automobiles, electricity, computers, CDs, and many other everyday conveniences.

Although chemistry is an ancient science, its modern foundation was laid in the nineteenth century, when intellectual and technological advances enabled scientists to break down substances into ever smaller components and consequently to explain many of their physical and chemical characteristics. The rapid development of increasingly sophisticated technology throughout the twentieth century has given us even greater means to study things that cannot be seen with the naked eye. Using computers and special microscopes, for example, chemists can analyze the structure of atoms and molecules—the fundamental units on which the study of chemistry is based—and design new substances with specific properties, such as drugs and environmentally friendly consumer products.

As we enter the twenty-first century, it is fitting to ask what part the central science will have in this century. Almost certainly, chemistry will continue to play a pivotal role in all areas of science and technology. Before plunging into the study of matter and its transformation, let us consider some of the frontiers that chemists are currently exploring (Figure 1.1). Whatever your reasons for taking general chemistry, a good knowledge of the subject will better enable you to appreciate its impact on society and on you as an individual.

Health and Medicine

Three major advances in the past century have enabled us to prevent and treat diseases. They are public health measures establishing sanitation systems to protect vast numbers of people from infectious disease; surgery with anesthesia, enabling physicians to cure potentially fatal conditions, such as an inflamed appendix; and the introduction of vaccines and antibiotics that make it possible to prevent diseases spread by microbes. Gene therapy promises to be the fourth revolution in medicine. (A gene is the basic unit of inheritance.) Several thousand known conditions, including cystic fibrosis and hemophilia, are carried by inborn damage to a single gene. Many other ailments, such as cancer, heart disease, AIDS, and arthritis, result to an extent from impairment of one or more genes involved in the body's defenses. In gene therapy, a selected healthy gene is delivered to a patient's cell to cure or ease such disorders. To carry out such a procedure, a doctor must have a sound knowledge of the chemical properties of the molecular components involved. The decoding of the human genome, which comprises all of the genetic material in the human body and plays an essential part in gene therapy, relies largely on chemical techniques.

Chemists in the pharmaceutical industry are researching potent drugs with few or no side effects to treat cancer, AIDS, and many other diseases as well as drugs to increase the number of successful organ transplants. On a broader scale, improved understanding of the mechanism of aging will lead to a longer and healthier life span for the world's population.

Energy and the Environment

Energy is a by-product of many chemical processes, and as the demand for energy continues to increase, both in technologically advanced countries like the United

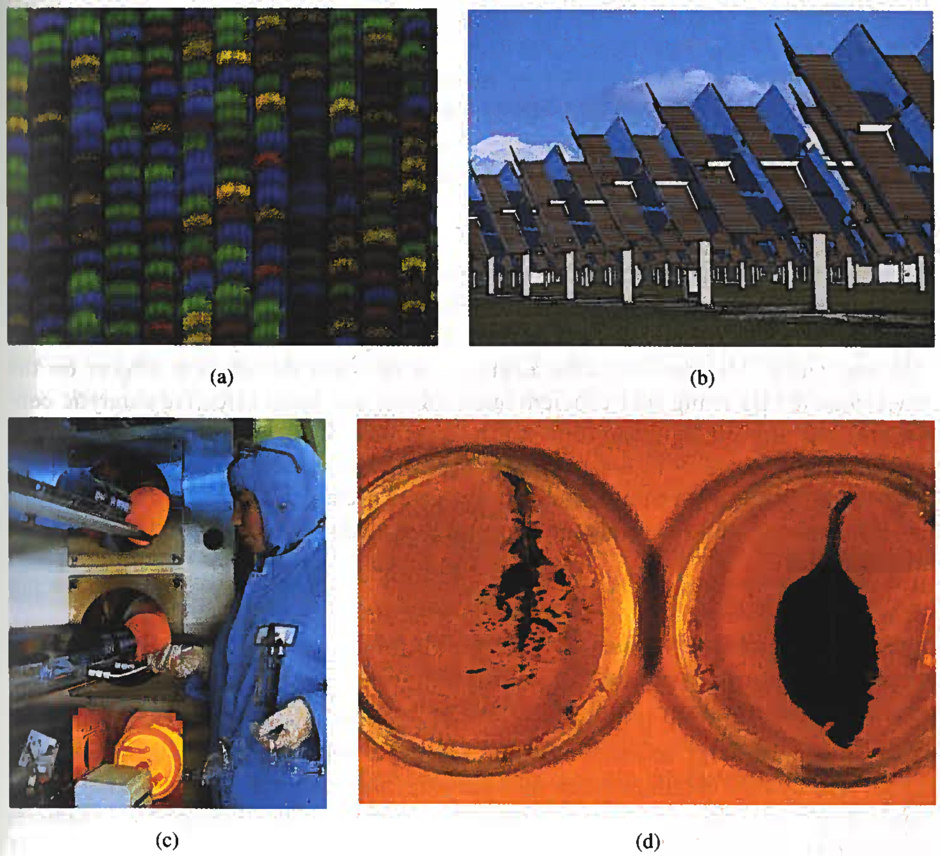


Figure 1.1 *(a) The output from an automated DNA sequencing machine. Each lane displays the sequence (indicated by different colors) obtained with a separate DNA sample. (b) Photovoltaic cells. (c) A silicon wafer being processed. (d) The leaf on the left was taken from a tobacco plant that was not genetically engineered but was exposed to tobacco horn worms. The leaf on the right was genetically engineered and is barely attacked by the worms. The same technique can be applied to protect the leaves of other types of plants.*

States and in developing ones like China, chemists are actively trying to find new energy sources. Currently the major sources of energy are fossil fuels (coal, petroleum, and natural gas). The estimated reserves of these fuels will last us another 50–100 years, at the present rate of consumption, so it is urgent that we find alternatives.

Solar energy promises to be a viable source of energy for the future. Every year Earth's surface receives about 10 times as much energy from sunlight as is contained in all of the known reserves of coal, oil, natural gas, and uranium combined. But much of this energy is "wasted" because it is reflected back into space. For the past 30 years, intense research efforts have shown that solar energy can be harnessed effectively in two ways. One is the conversion of sunlight directly to electricity using devices called *photovoltaic cells*. The other is to use sunlight to obtain hydrogen from water. The hydrogen can then be fed into a *fuel cell* to generate electricity. Although our understanding of the scientific process of converting solar energy to electricity has advanced, the technology has not yet improved to the point where we can produce electricity on a large scale at an economically acceptable cost. By 2050, however, it has been predicted that solar energy will supply over 50 percent of our power needs.

Another potential source of energy is nuclear fission, but because of environmental concerns about the radioactive wastes from fission processes, the future of the nuclear industry in the United States is uncertain. Chemists can help to devise better ways to dispose of nuclear waste. Nuclear fusion, the process that occurs in the sun and other stars, generates huge amounts of energy without producing much dangerous radioactive waste. In another 50 years, nuclear fusion will likely be a significant source of energy.

Energy production and energy utilization are closely tied to the quality of our environment. A major disadvantage of burning fossil fuels is that they give off carbon dioxide, which is a *greenhouse gas* (that is, it promotes the heating of Earth's atmosphere), along with sulfur dioxide and nitrogen oxides, which result in acid rain and smog. (Harnessing solar energy has no such detrimental effects on the environment.) By using fuel-efficient automobiles and more effective catalytic converters, we should be able to drastically reduce harmful auto emissions and improve the air quality in areas with heavy traffic. In addition, electric cars, powered by durable, long-lasting batteries, and hybrid cars, powered by both batteries and gasoline, should become more prevalent, and their use will help to minimize air pollution.

Materials and Technology

Chemical research and development in the twentieth century have provided us with new materials that have profoundly improved the quality of our lives and helped to advance technology in countless ways. A few examples are polymers (including rubber and nylon), ceramics (such as cookware), liquid crystals (like those in electronic displays), adhesives (used in your Post-It notes), and coatings (for example, latex paint).

What is in store for the near future? One likely possibility is room-temperature *superconductors.* Electricity is carried by copper cables, which are not perfect conductors. Consequently, about 20 percent of electrical energy is lost in the form of heat between the power station and our homes. This is a tremendous waste. Superconductors are materials that have no electrical resistance and can therefore conduct electricity with no energy loss. Although the phenomenon of superconductivity at very low temperatures (more than 400 degrees Fahrenheit below the freezing point of water) has been known for over 80 years, a major breakthrough in the mid-1980s demonstrated that it is possible to make materials that act as superconductors at or near room temperature. Chemists have helped to design and synthesize new materials that show promise in this quest. The next 30 years will see high-temperature superconductors being applied on a large scale in magnetic resonance imaging (MRI), levitated trains, and nuclear fusion.

If we had to name one technological advance that has shaped our lives more than any other, it would be the computer. The "engine" that drives the ongoing computer revolution is the microprocessor—the tiny silicon chip that has inspired countless inventions, such as laptop computers and fax machines. The performance of a microprocessor is judged by the speed with which it carries out mathematical operations, such as addition. The pace of progress is such that since their introduction, microprocessors have doubled in speed every 18 months. The quality of any microprocessor depends on the purity of the silicon chip and on the ability to add the desired amount of other substances, and chemists play an important role in the research and development of silicon chips. For the future, scientists have begun to explore the prospect of "molecular computing," that is, replacing silicon with molecules. The advantages are that certain molecules can be made to respond to light, rather than to electrons, so that we would have optical computers rather than electronic computers. With proper genetic engineering, scientists can synthesize such molecules using microorganisms instead of large factories. Optical computers also would have much greater storage capacity than electronic computers.

Food and Agriculture

How can the world's rapidly increasing population be fed? In poor countries, agricultural activities occupy about 80 percent of the workforce, and half of an average family budget is spent on foodstuffs. This is a tremendous drain on a nation's resources. The factors that affect agricultural production are the richness of the soil, insects and diseases that damage crops, and weeds that compete for nutrients. Besides irrigation, farmers rely on fertilizers and pesticides to increase crop yield. Since the 1950s, treatment for crops suffering from pest infestations has sometimes been the indiscriminate application of potent chemicals. Such measures have often had serious detrimental effects on the environment. Even the excessive use of fertilizers is harmful to the land, water, and air.

To meet the food demands of the twenty-first century, new and novel approaches in farming must be devised. It has already been demonstrated that, through biotechnology, it is possible to grow larger and better crops. These techniques can be applied to many different farm products, not only for improved yields, but also for better frequency, that is, more crops every year. For example, it is known that a certain bacterium produces a protein molecule that is toxic to leaf-eating caterpillars. Incorporating the gene that codes for the toxin into crops enables plants to protect themselves so that pesticides are not necessary. Researchers have also found a way to prevent pesky insects from reproducing. Insects communicate with one another by emitting and reacting to special molecules called pheromones. By identifying and synthesizing pheromones used in mating, it is possible to interfere with the normal reproductive cycle of common pests; for example, by inducing insects to mate too soon or tricking female insects into mating with sterile males. Moreover, chemists can devise ways to increase the production of fertilizers that are less harmful to the environment and substances that would selectively kill weeds.

1.2 The Study of Chemistry

Compared with other subjects, chemistry is commonly believed to be more difficult, at least at the introductory level. There is some justification for this perception; for one thing, chemistry has a very specialized vocabulary. However, even if this is your first course in chemistry, you already have more familiarity with the subject than you may realize. In everyday conversations we hear words that have a chemical connection, although they may not be used in the scientifically correct sense. Examples are "electronic," "quantum leap," "equilibrium," "catalyst," "chain reaction," and "critical mass." Moreover, if you cook, then you are a practicing chemist! From experience gained in the kitchen, you know that oil and water do not mix and that boiling water left on the stove will evaporate. You apply chemical and physical principles when you use baking soda to leaven bread, choose a pressure cooker to shorten the time it takes to prepare soup, add meat tenderizer to a pot roast, squeeze lemon juice over sliced pears to prevent them from turning brown or over fish to minimize its odor, and add vinegar to the water in which you are going to poach eggs. Every day we observe such changes without thinking about their chemical nature. The purpose of this course

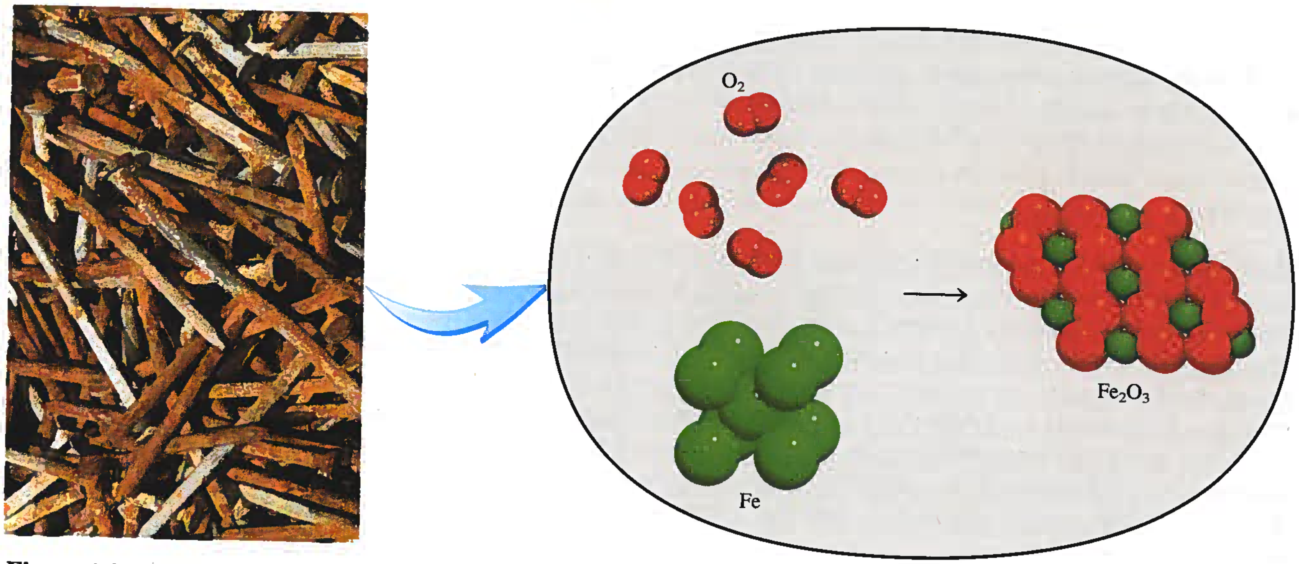


Figure 1.2 *A simplified molecular view of rust (Fe_2O_3) formation from iron atoms (Fe) and oxygen molecules (O_2). In reality, the process requires water and rust also contains water molecules.*

is to make you think like a chemist, to look at the *macroscopic world*—the things we can see, touch, and measure directly—and visualize the particles and events of the *microscopic world* that we cannot experience without modern technology and our imaginations.

At first some students find it confusing that their chemistry instructor and textbook seem to be continually shifting back and forth between the macroscopic and microscopic worlds. Just keep in mind that the data for chemical investigations most often come from observations of large-scale phenomena, but the explanations frequently lie in the unseen and partially imagined microscopic world of atoms and molecules. In other words, chemists often *see* one thing (in the macroscopic world) and *think* another (in the microscopic world). Looking at the rusted nails in Figure 1.2, for example, a chemist might think about the basic properties of individual atoms of iron and how these units interact with other atoms and molecules to produce the observed change.

1.3 The Scientific Method

All sciences, including the social sciences, employ variations of what is called the ***scientific method,*** *a systematic approach to research.* For example, a psychologist who wants to know how noise affects people's ability to learn chemistry and a chemist interested in measuring the heat given off when hydrogen gas burns in air would follow roughly the same procedure in carrying out their investigations. The first step is to carefully define the problem. The next step includes performing experiments, making careful observations, and recording information, or *data,* about the system—the part of the universe that is under investigation. (In the examples just discussed, the systems are the group of people the psychologist will study and a mixture of hydrogen and air.)

The data obtained in a research study may be both ***qualitative,*** *consisting of general observations about the system,* and ***quantitative,*** *comprising numbers obtained*

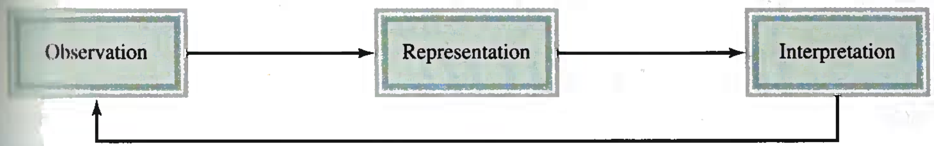


Figure 1.3 *The three levels of studying chemistry and their relationships. Observation deals with events in the macroscopic world; atoms and molecules constitute the microscopic world. Representation is a scientific shorthand for describing an experiment in symbols and chemical equations. Chemists use their knowledge of atoms and molecules to explain an observed phenomenon.*

by various measurements of the system. Chemists generally use standardized symbols and equations in recording their measurements and observations. This form of representation not only simplifies the process of keeping records, but also provides a common basis for communication with other chemists.

When the experiments have been completed and the data have been recorded, the next step in the scientific method is interpretation, meaning that the scientist attempts to explain the observed phenomenon. Based on the data that were gathered, the researcher formulates a ***hypothesis,*** *a tentative explanation for a set of observations.* Further experiments are devised to test the validity of the hypothesis in as many ways as possible, and the process begins anew. Figure 1.3 summarizes the main steps of the research process.

After a large amount of data has been collected, it is often desirable to summarize the information in a concise way, as a law. In science, a ***law*** is *a concise verbal or mathematical statement of a relationship between phenomena that is always the same under the same conditions.* For example, Sir Isaac Newton's second law of motion, which you may remember from high school science, says that force equals mass times acceleration ($F = ma$). What this law means is that an increase in the mass or in the acceleration of an object will always increase its force proportionally, and a decrease in mass or acceleration will always decrease the force.

Hypotheses that survive many experimental tests of their validity may evolve into theories. A ***theory*** is *a unifying principle that explains a body of facts and/or those laws that are based on them.* Theories, too, are constantly being tested. If a theory is disproved by experiment, then it must be discarded or modified so that it becomes consistent with experimental observations. Proving or disproving a theory can take years, even centuries, in part because the necessary technology may not be available. Atomic theory, which we will study in Chapter 2, is a case in point. It took more than 2000 years to work out this fundamental principle of chemistry proposed by Democritus, an ancient Greek philosopher. A more contemporary example is the Big Bang theory of the origin of the universe discussed on page 10.

Scientific progress is seldom, if ever, made in a rigid, step-by-step fashion. Sometimes a law precedes a theory; sometimes it is the other way around. Two scientists may start working on a project with exactly the same objective, but will end up taking drastically different approaches. Scientists are, after all, human beings, and their modes of thinking and working are very much influenced by their background, training, and personalities.

The development of science has been irregular and sometimes even illogical. Great discoveries are usually the result of the cumulative contributions and experience of many workers, even though the credit for formulating a theory or a law is usually given to only one individual. There is, of course, an element of luck involved in scientific discoveries, but it has been said that "chance favors the prepared mind." It takes an alert and well-trained person to recognize the significance of an accidental discovery and to take full advantage of it. More often than not, the public learns only of spectacular scientific breakthroughs. For every success story, however, there are hundreds of cases in which scientists have spent years working on projects that

CHEMISTRY *in Action*

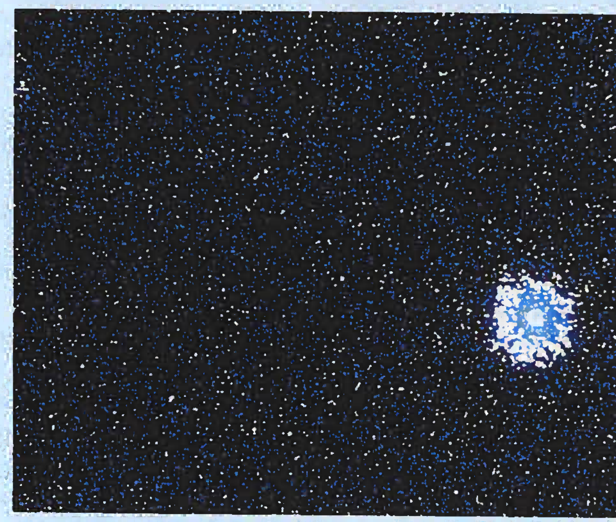
Primordial Helium and the Big Bang Theory

Where did we come from? How did the universe begin? Humans have asked these questions for as long as we have been able to think. The search for answers provides an example of the scientific method.

In the 1940s the Russian-American physicist George Gamow hypothesized that our universe burst into being billions of years ago in a gigantic explosion, or *Big Bang*. In its earliest moments, the universe occupied a tiny volume and was unimaginably hot. This blistering fireball of radiation mixed with microscopic particles of matter gradually cooled enough for atoms to form. Under the influence of gravity, these atoms clumped together to make billions of galaxies including our own Milky Way Galaxy.

Gamow's idea is interesting and highly provocative. It has been tested experimentally in a number of ways. First, measurements showed that the universe is expanding; that is, galaxies are all moving away from one another at high speeds. This fact is consistent with the universe's explosive birth. By imagining the expansion running backward, like a movie in reverse, astronomers have deduced that the universe was born about 13 billion years ago. The second observation that supports Gamow's hypothesis is the detection of *cosmic background radiation.* Over billions of years, the searingly hot universe has cooled down to a mere 3 K (or −270°C)! At this temperature, most energy is in the microwave region. Because the Big Bang would have occurred simultaneously throughout the tiny volume of the forming universe, the radiation it generated should have filled the entire universe. Thus, the radiation should be the same in any direction that we observe. Indeed, the microwave signals recorded by astronomers are *independent* of direction.

The third piece of evidence supporting Gamow's hypothesis is the discovery of primordial helium. Scientists believe that helium and hydrogen (the lightest elements) were the first elements formed in the early stages of cosmic evolution. (The heavier elements, like carbon, nitrogen, and oxygen, are thought to have originated later via nuclear reactions involving hydrogen and helium in the center of stars.) If so, a diffuse gas of hydrogen and helium would have spread through the early universe before much of the galaxies formed. In 1995



A color photo of some distant galaxy, including the position of a quasar.

astronomers analyzed ultraviolet light from a distant *quasar* (a strong source of light and radio signals that is thought to be an exploding galaxy at the edge of the universe) and found that some of the light was absorbed by helium atoms on the way to Earth. Because this particular quasar is more than 10 billion light-years away (a light-year is the distance traveled by light in a year), the light reaching Earth reveals events that took place 10 billion years ago. Why wasn't the more abundant hydrogen detected? A hydrogen atom has only one electron, which is stripped by the light from a quasar in a process known as *ionization.* Ionized hydrogen atoms cannot absorb any of the quasar's light. A helium atom, on the other hand, has two electrons. Radiation may strip a helium atom of one electron, but not always both. Singly ionized helium atoms can still absorb light and are therefore detectable.

Proponents of Gamow's explanation rejoiced at the detection of helium in the far reaches of the universe. In recognition of all the supporting evidence, scientists now refer to Gamow's hypothesis as the Big Bang theory.

ultimately led to a dead end, and in which positive achievements came only after many wrong turns and at such a slow pace that they went unheralded. Yet even the dead ends contribute something to the continually growing body of knowledge about the physical universe. It is the love of the search that keeps many scientists in the laboratory.

1.4 Classifications of Matter

We defined chemistry at the beginning of the chapter as the study of matter and the changes it undergoes. ***Matter*** is *anything that occupies space and has mass.* Matter includes things we can see and touch (such as water, earth, and trees), as well as things we cannot (such as air). Thus, everything in the universe has a "chemical" connection.

Chemists distinguish among several subcategories of matter based on composition and properties. The classifications of matter include substances, mixtures, elements, and compounds, as well as atoms and molecules, which we will consider in Chapter 2.

Substances and Mixtures

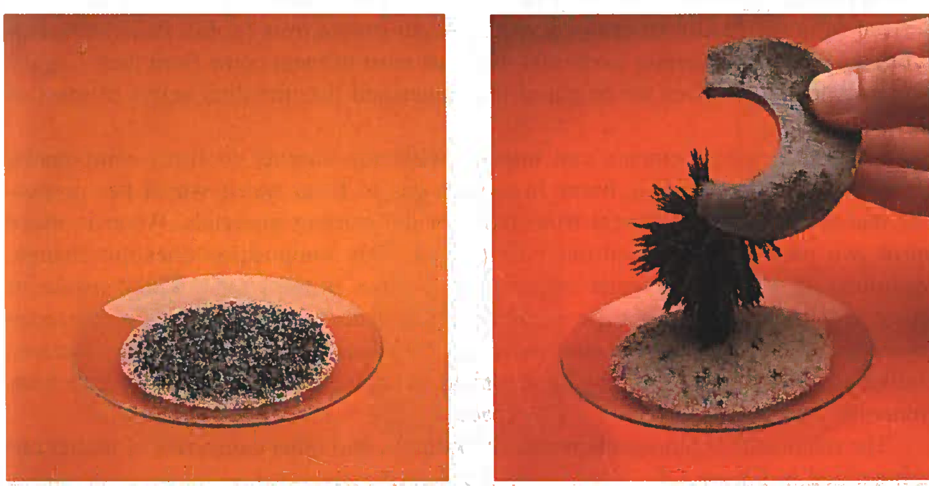
Interactivity:
Substances and Mixtures
ARIS, Interactives

A ***substance*** is *a form of matter that has a definite (constant) composition and distinct properties.* Examples are water, ammonia, table sugar (sucrose), gold, and oxygen. Substances differ from one another in composition and can be identified by their appearance, smell, taste, and other properties.

A ***mixture*** is *a combination of two or more substances in which the substances retain their distinct identities.* Some familiar examples are air, soft drinks, milk, and cement. Mixtures do not have constant composition. Therefore, samples of air collected in different cities would probably differ in composition because of differences in altitude, pollution, and so on.

Mixtures are either homogeneous or heterogeneous. When a spoonful of sugar dissolves in water we obtain a ***homogeneous mixture*** in which *the composition of the mixture is the same throughout.* If sand is mixed with iron filings, however, the sand grains and the iron filings remain separate (Figure 1.4). This type of mixture is called a ***heterogeneous mixture*** because *the composition is not uniform.*

Any mixture, whether homogeneous or heterogeneous, can be created and then separated by physical means into pure components without changing the identities of the components. Thus, sugar can be recovered from a water solution by heating the solution and evaporating it to dryness. Condensing the vapor will give us back the water component. To separate the iron-sand mixture, we can use a magnet to remove



(a) (b)

Figure 1.4 *(a) The mixture contains iron filings and sand. (b) A magnet separates the iron filings from the mixture. The same technique is used on a larger scale to separate iron and steel from nonmagnetic objects such as aluminum, glass, and plastics.*

TABLE 1.1 Some Common Elements and Their Symbols

Name	Symbol	Name	Symbol	Name	Symbol
Aluminum	Al	Fluorine	F	Oxygen	O
Arsenic	As	Gold	Au	Phosphorus	P
Barium	Ba	Hydrogen	H	Platinum	Pt
Bismuth	Bi	Iodine	I	Potassium	K
Bromine	Br	Iron	Fe	Silicon	Si
Calcium	Ca	Lead	Pb	Silver	Ag
Carbon	C	Magnesium	Mg	Sodium	Na
Chlorine	Cl	Manganese	Mn	Sulfur	S
Chromium	Cr	Mercury	Hg	Tin	Sn
Cobalt	Co	Nickel	Ni	Tungsten	W
Copper	Cu	Nitrogen	N	Zinc	Zn

the iron filings from the sand, because sand is not attracted to the magnet [see Figure 1.4(b)]. After separation, the components of the mixture will have the same composition and properties as they did to start with.

Elements and Compounds

Interactivity:
Elements
ARIS, Interactives

Substances can be either elements or compounds. An ***element*** is *a substance that cannot be separated into simpler substances by chemical means.* To date, 114 elements have been positively identified. Most of them occur naturally on Earth. The others have been created by scientists via nuclear processes, which are the subject of Chapter 23 of this text.

For convenience, chemists use symbols of one or two letters to represent the elements. The first letter of a symbol is *always* capitalized, but any following letters are not. For example, Co is the symbol for the element cobalt, whereas CO is the formula for the carbon monoxide molecule. Table 1.1 shows the names and symbols of some of the more common elements; a complete list of the elements and their symbols appears inside the front cover of this book. The symbols of some elements are derived from their Latin names—for example, Au from *aurum* (gold), Fe from *ferrum* (iron), and Na from *natrium* (sodium)—whereas most of them come from their English names. Appendix 1 gives the origin of the names and lists the discoverers of most of the elements.

Atoms of most elements can interact with one another to form compounds. Hydrogen gas, for example, burns in oxygen gas to form water, which has properties that are distinctly different from those of the starting materials. Water is made up of two parts hydrogen and one part oxygen. This composition does not change, regardless of whether the water comes from a faucet in the United States, a lake in Outer Mongolia, or the ice caps on Mars. Thus, water is a ***compound,*** *a substance composed of atoms of two or more elements chemically united in fixed proportions.* Unlike mixtures, compounds can be separated only by chemical means into their pure components.

The relationships among elements, compounds, and other categories of matter are summarized in Figure 1.5.

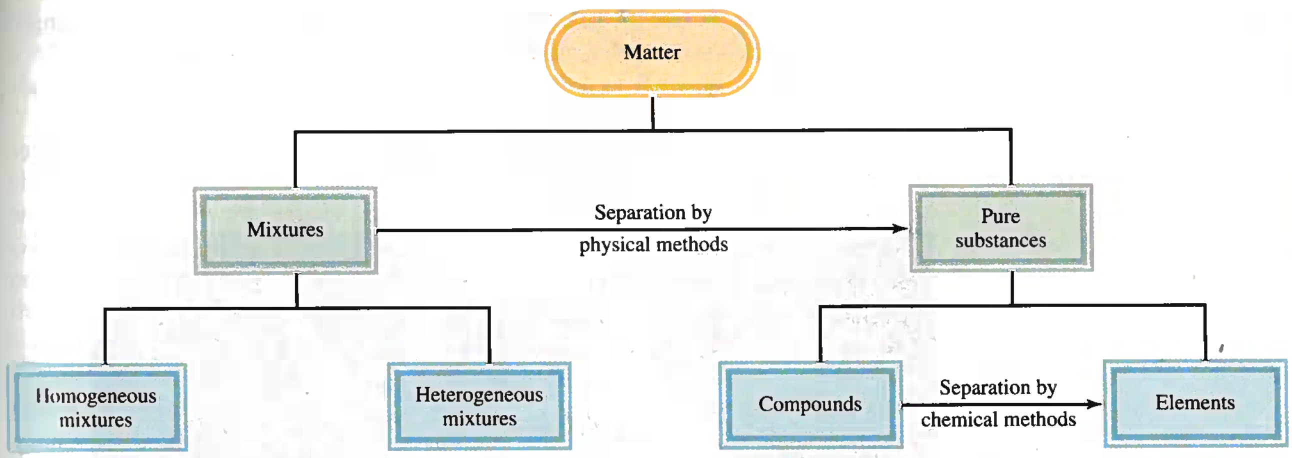


Figure 1.5 *Classifications of matter.*

1.5 The Three States of Matter

All substances, at least in principle, can exist in three states: solid, liquid, and gas. As Figure 1.6 shows, gases differ from liquids and solids in the distances between the molecules. In a solid, molecules are held close together in an orderly fashion with little freedom of motion. Molecules in a liquid are close together but are not held so rigidly in position and can move past one another. In a gas, the molecules are separated by distances that are large compared with the size of the molecules.

The three states of matter can be interconverted without changing the composition of the substance. Upon heating, a solid (for example, ice) will melt to form a liquid (water). (The temperature at which this transition occurs is called the *melting point.*) Further heating will convert the liquid into a gas. (This conversion takes place at the *boiling point* of the liquid.) On the other hand, cooling a gas will cause it to condense into a liquid. When the liquid is cooled further, it will freeze into the solid

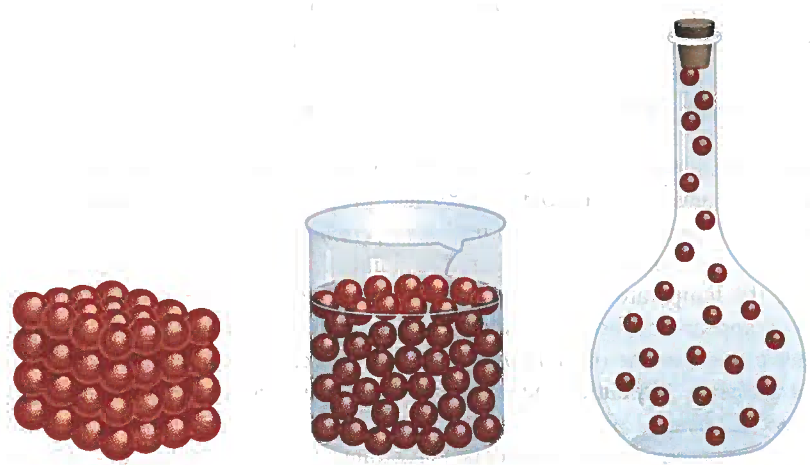


Figure 1.6 *Microscopic views of a solid, a liquid, and a gas.*

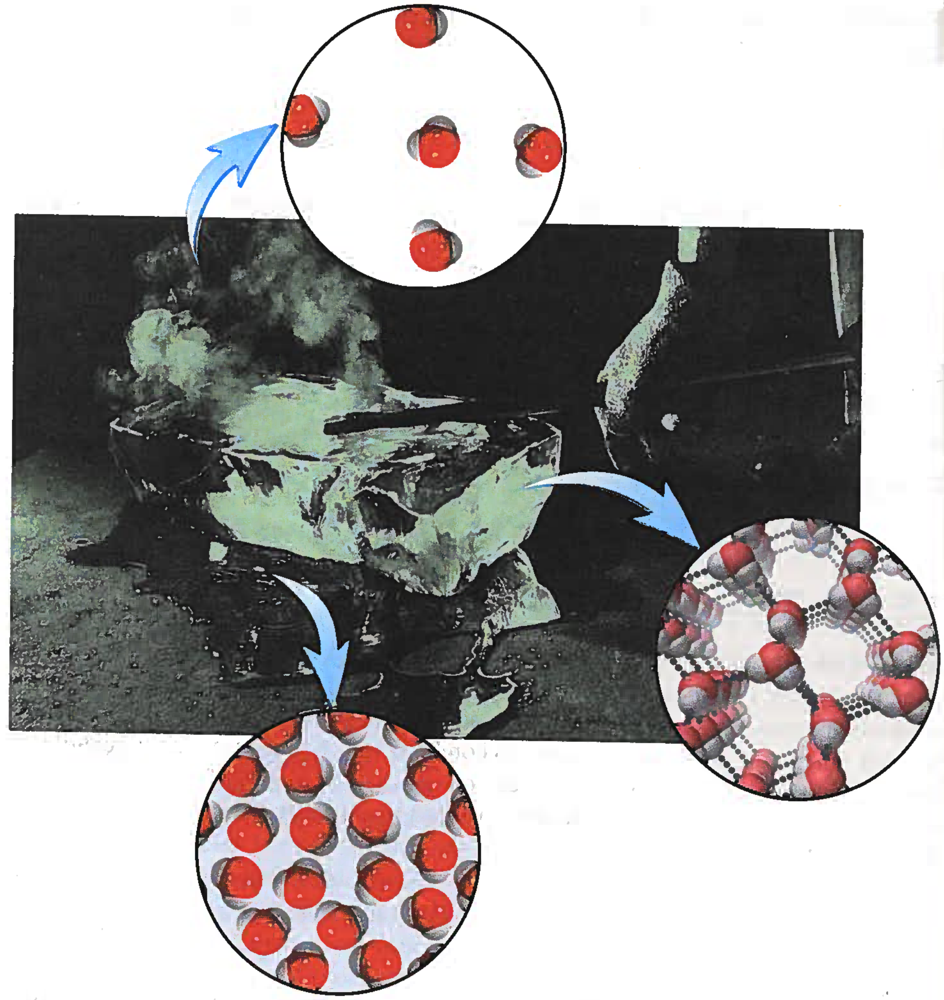


Figure 1.7 *The three states of matter. A hot poker changes ice into water and steam.*

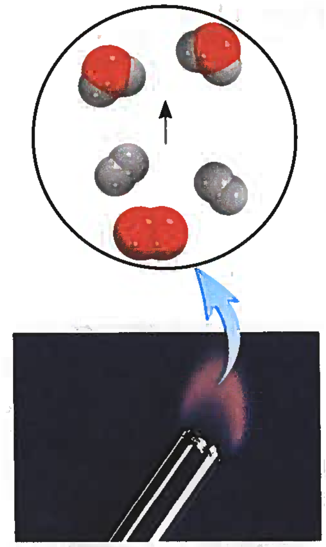
form. Figure 1.7 shows the three states of water. Note that the properties of water are unique among common substances in that the molecules in the liquid state are more closely packed than those in the solid state.

1.6 Physical and Chemical Properties of Matter

Substances are identified by their properties as well as by their composition. Color, melting point, and boiling point are physical properties. A ***physical property*** *can be measured and observed without changing the composition or identity of a substance.* For example, we can measure the melting point of ice by heating a block of ice and recording the temperature at which the ice is converted to water. Water differs from ice only in appearance, not in composition, so this is a physical change; we can freeze the water to recover the original ice. Therefore, the melting point of a substance is a physical property. Similarly, when we say that helium gas is lighter than air, we are referring to a physical property.

On the other hand, the statement "Hydrogen gas burns in oxygen gas to form water" describes a ***chemical property*** of hydrogen, because *to observe this property we must carry out a chemical change,* in this case burning. After the change, the original chemical substance, the hydrogen gas, will have vanished, and all that will be left is a different chemical substance—water. We *cannot* recover the hydrogen from the water by means of a physical change, such as boiling or freezing.

Every time we hard-boil an egg, we bring about a chemical change. When subjected to a temperature of about 100°C, the yolk and the egg white undergo changes that alter not only their physical appearance but their chemical makeup as well. When eaten, the egg is changed again, by substances in our bodies called *enzymes*. This digestive action is another example of a chemical change. What happens during digestion depends on the chemical properties of both the enzymes and the food.



Hydrogen burning in air to form water.

All measurable properties of matter fall into one of two additional categories: extensive properties and intensive properties. The measured value of an ***extensive property*** *depends on how much matter is being considered.* ***Mass,*** which is *the quantity of matter in a given sample of a substance,* is an extensive property. More matter means more mass. Values of the same extensive property can be added together. For example, two copper pennies will have a combined mass that is the sum of the masses of each penny, and the length of two tennis courts is the sum of the lengths of each tennis court. ***Volume,*** defined as *length cubed,* is another extensive property. The value of an extensive quantity depends on the amount of matter.

The measured value of an ***intensive property*** *does not depend on how much matter is being considered.* ***Density,*** defined as *the mass of an object divided by its volume,* is an intensive property. So is temperature. Suppose that we have two beakers of water at the same temperature. If we combine them to make a single quantity of water in a larger beaker, the temperature of the larger quantity of water will be the same as it was in two separate beakers. Unlike mass, length, and volume, temperature and other intensive properties are not additive.

1.7 Measurement

The measurements chemists make are often used in calculations to obtain other related quantities. Different instruments enable us to measure a substance's properties: The meterstick measures length or scale; the buret, the pipet, the graduated cylinder, and the volumetric flask measure volume (Figure 1.8); the balance measures mass; the thermometer measures temperature. These instruments provide measurements of ***macroscopic properties,*** which *can be determined directly.* ***Microscopic properties,*** *on the atomic or molecular scale, must be determined by an indirect method,* as we will see in Chapter 2.

A measured quantity is usually written as a number with an appropriate unit. To say that the distance between New York and San Francisco by car along a certain route is 5166 is meaningless. We must specify that the distance is 5166 kilometers. The same is true in chemistry; units are essential to stating measurements correctly.

SI Units

For many years scientists recorded measurements in *metric units,* which are related decimally, that is, by powers of 10. In 1960, however, the General Conference of Weights and Measures, the international authority on units, proposed a revised metric system called the ***International System of Units*** (abbreviated ***SI,*** from the French *S*ystème *I*nternationale d'Unites). Table 1.2 shows the seven SI base units. All

Interactivity:
SI Base Units
ARIS, Interactives

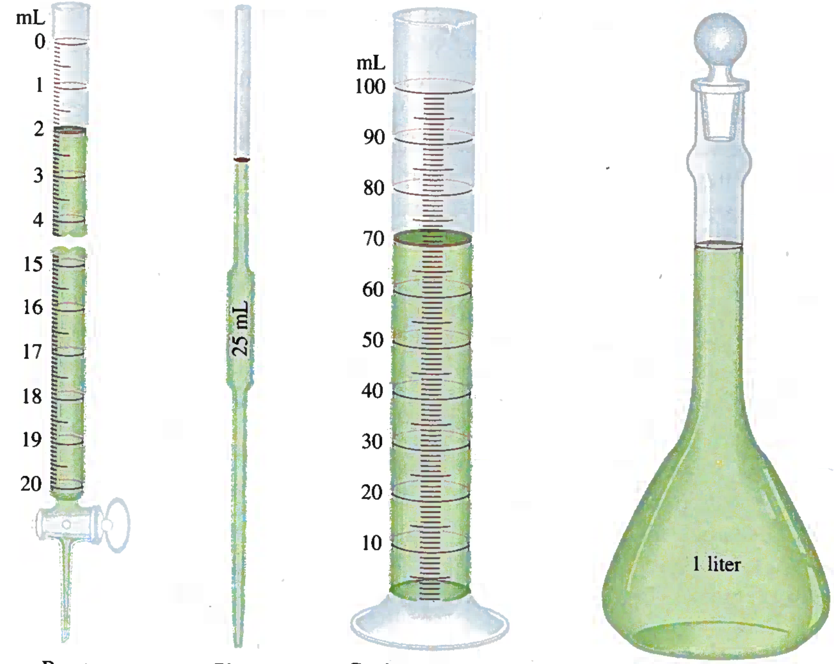
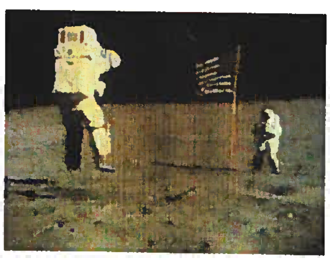


Figure 1.8 *Some common measuring devices found in a chemistry laboratory. These devices are not drawn to scale relative to one another. We will discuss the uses of these measuring devices in Chapter 4.*

Interactivity:
Unit Prefixes
ARIS, Interactives

other units of measurement can be derived from these base units. Like metric units, SI units are modified in decimal fashion by a series of prefixes, as shown in Table 1.3. We will use both metric and SI units in this book.

Measurements that we will utilize frequently in our study of chemistry include time, mass, volume, density, and temperature.



An astronaut jumping on the surface of the moon.

Mass and Weight

The terms "mass" and "weight" are often used interchangeably, although, strictly speaking, they are different quantities. Whereas mass is a measure of the amount of matter in an object, ***weight,*** technically speaking, is *the force that gravity exerts on an object.* An apple that falls from a tree is pulled downward by Earth's gravity. The mass of the apple is constant and does not depend on its location, but its weight does. For example, on the surface of the moon the apple would weigh only one-sixth what it does on Earth, because

TABLE 1.2 SI Base Units

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

TABLE 1.3 Prefixes Used with SI Units

Prefix	Symbol	Meaning	Example
tera-	T	1,000,000,000,000, or 10^{12}	1 terameter (Tm) $= 1 \times 10^{12}$ m
giga-	G	1,000,000,000, or 10^{9}	1 gigameter (Gm) $= 1 \times 10^{9}$ m
mega-	M	1,000,000, or 10^{6}	1 megameter (Mm) $= 1 \times 10^{6}$ m
kilo-	k	1,000, or 10^{3}	1 kilometer (km) $= 1 \times 10^{3}$ m
deci-	d	1/10, or 10^{-1}	1 decimeter (dm) = 0.1 m
centi-	c	1/100, or 10^{-2}	1 centimeter (cm) = 0.01 m
milli-	m	1/1,000, or 10^{-3}	1 millimeter (mm) = 0.001 m
micro-	μ	1/1,000,000, or 10^{-6}	1 micrometer (μm) $= 1 \times 10^{-6}$ m
nano-	n	1/1,000,000,000, or 10^{-9}	1 nanometer (nm) $= 1 \times 10^{-9}$ m
pico-	p	1/1,000,000,000,000, or 10^{-12}	1 picometer (pm) $= 1 \times 10^{-12}$ m

the moon's gravity is only one-sixth that of Earth. The moon's smaller gravity enabled astronauts to jump about rather freely on its surface despite their bulky suits and equipment. Chemists are interested primarily in mass, which can be determined readily with a balance; the process of measuring mass, oddly, is called *weighing.*

The SI unit of mass is the *kilogram* (kg). Unlike the units of length and time, which are based on natural processes that can be repeated by scientists anywhere, the kilogram is defined in terms of a particular object (Figure 1.9). In chemistry, however, the smaller *gram* (g) is more convenient:

$$1 \text{ kg} = 1000 \text{ g} = 1 \times 10^{3} \text{ g}$$

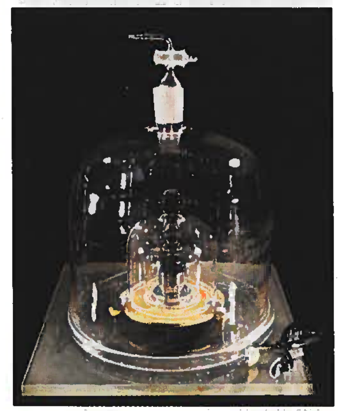


Figure 1.9 *The prototype kilogram is made of a platinum-iridium alloy. It is kept in a vault at the International Bureau of Weights and Measures in Sèvres, France.*

Volume

The SI unit of length is the *meter* (m), and the SI-derived unit for volume is the *cubic meter* (m^3). Generally, however, chemists work with much smaller volumes, such as the cubic centimeter (cm^3) and the cubic decimeter (dm^3):

$$1 \text{ cm}^3 = (1 \times 10^{-2} \text{ m})^3 = 1 \times 10^{-6} \text{ m}^3$$
$$1 \text{ dm}^3 = (1 \times 10^{-1} \text{ m})^3 = 1 \times 10^{-3} \text{ m}^3$$

Another common unit of volume is the liter (L). A ***liter*** is *the volume occupied by one cubic decimeter.* One liter of volume is equal to 1000 milliliters (mL) or 1000 cm^3:

$$\begin{aligned} 1 \text{ L} &= 1000 \text{ mL} \\ &= 1000 \text{ cm}^3 \\ &= 1 \text{ dm}^3 \end{aligned}$$

and one milliliter is equal to one cubic centimeter:

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

Figure 1.10 compares the relative sizes of two volumes. Even though the liter is not an SI unit, volumes are usually expressed in liters and milliliters.

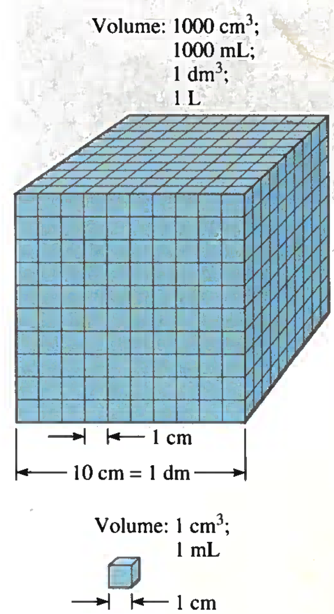


Figure 1.10 *Comparison of two volumes, 1 mL and 1000 mL.*

Interactivity:
Density
ARIS, Interactives

Density

The equation for density is

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

or

$$d = \frac{m}{V} \tag{1.1}$$

where d, m, and V denote density, mass, and volume, respectively. Because density is an intensive property and does not depend on the quantity of mass present, for a given substance the ratio of mass to volume always remains the same; in other words, V increases as m does.

The SI-derived unit for density is the kilogram per cubic meter (kg/m^3). This unit is awkwardly large for most chemical applications. Therefore, grams per cubic centimeter (g/cm^3) and its equivalent, grams per milliliter (g/mL), are more commonly used for solid and liquid densities. Because gas densities are often very low, we express them in units of grams per liter (g/L):

$$1 \text{ g/cm}^3 = 1 \text{ g/mL} = 1000 \text{ kg/m}^3$$
$$1 \text{ g/L} = 0.001 \text{ g/mL}$$

Table 1.4 lists the densities of several substances.

Examples 1.1 and 1.2 show density calculations.

TABLE 1.4
Densities of Some Substances at 25°C

Substance	Density (g/cm^3)
Air*	0.001
Ethanol	0.79
Water	1.00
Mercury	13.6
Table salt	2.2
Iron	7.9
Gold	19.3
Osmium†	22.6

*Measured at 1 atmosphere.
†Osmium (Os) is the densest element known.

Gold bars.

Example 1.1

Gold is a precious metal that is chemically unreactive. It is used mainly in jewelry, dentistry, and electronic devices. A piece of gold ingot with a mass of 301 g has a volume of 15.6 cm^3. Calculate the density of gold.

Solution We are given the mass and volume and asked to calculate the density. Therefore, from Equation (1.1), we write

$$d = \frac{m}{V}$$
$$= \frac{301 \text{ g}}{15.6 \text{ cm}^3}$$
$$= \boxed{19.3 \text{ g/cm}^3}$$

Similar problems: 1.21, 1.22.

Practice Exercise A piece of platinum metal with a density of 21.5 g/cm^3 has a volume of 4.49 cm^3. What is its mass?

Example 1.2

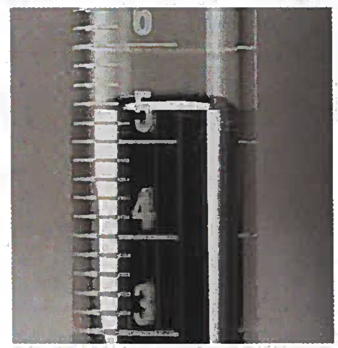
The density of mercury, the only metal that is a liquid at room temperature, is 13.6 g/mL. Calculate the mass of 5.50 mL of the liquid.

(Continued)

Solution We are given the density and volume of a liquid and asked to calculate the mass of the liquid. We rearrange Equation (1.1) to give

$$m = d \times V$$
$$= 13.6\frac{\text{g}}{\cancel{\text{mL}}} \times 5.50\ \cancel{\text{mL}}$$
$$= \boxed{74.8\ \text{g}}$$

Practice Exercise The density of sulfuric acid in a certain car battery is 1.41 g/mL. Calculate the mass of 242 mL of the liquid.



Mercury.

Similar problems: 1.21, 1.22.

Temperature Scales

Three temperature scales are currently in use. Their units are °F (degrees Fahrenheit), °C (degrees Celsius), and K (kelvin).The Fahrenheit scale, which is the most commonly used scale in the United States outside the laboratory, defines the normal freezing and boiling points of water to be exactly 32°F and 212°F, respectively. The Celsius scale divides the range between the freezing point (0°C) and boiling point (100°C) of water into 100 degrees. As Table 1.2 shows, the ***kelvin*** is *the SI base unit of temperature:* it is the *absolute* temperature scale. By absolute we mean that the zero on the Kelvin scale, denoted by 0 K, is the lowest temperature that can be attained theoretically. On the other hand, 0°F and 0°C are based on the behavior of an arbitrarily chosen substance, water. Figure 1.11 compares the three temperature scales.

Note that the Kelvin scale does not have the degree sign. Also, temperatures expressed in kelvins can never be negative.

The size of a degree on the Fahrenheit scale is only 100/180, or 5/9, of a degree on the Celsius scale. To convert degrees Fahrenheit to degrees Celsius, we write

$$?°\text{C} = (°\text{F} - 32°\text{F}) \times \frac{5°\text{C}}{9°\text{F}} \qquad (1.2)$$

The following equation is used to convert degrees Celsius to degrees Fahrenheit:

$$?°\text{F} = \frac{9°\text{F}}{5°\text{C}} \times (°\text{C}) + 32°\text{F} \qquad (1.3)$$

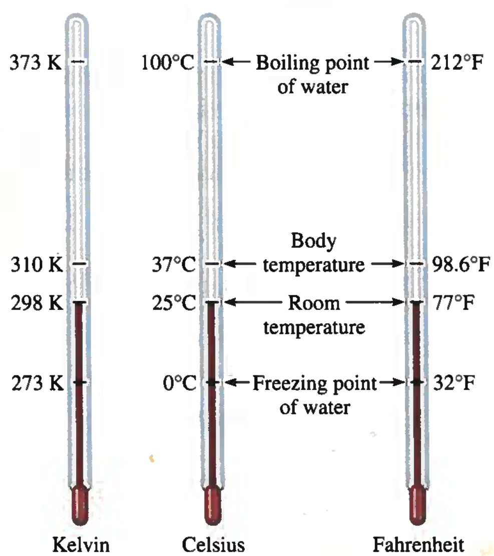


Figure 1.11 *Comparison of the three temperature scales: Celsius, Fahrenheit, and the absolute (Kelvin) scales. Note that there are 100 divisions, or 100 degrees, between the freezing point and the boiling point of water on the Celsius scale, and there are 180 divisions, or 180 degrees, between the same two temperature limits on the Fahrenheit scale. The Celsius scale was formerly called the centigrade scale.*

CHEMISTRY *in Action*

The Importance of Units

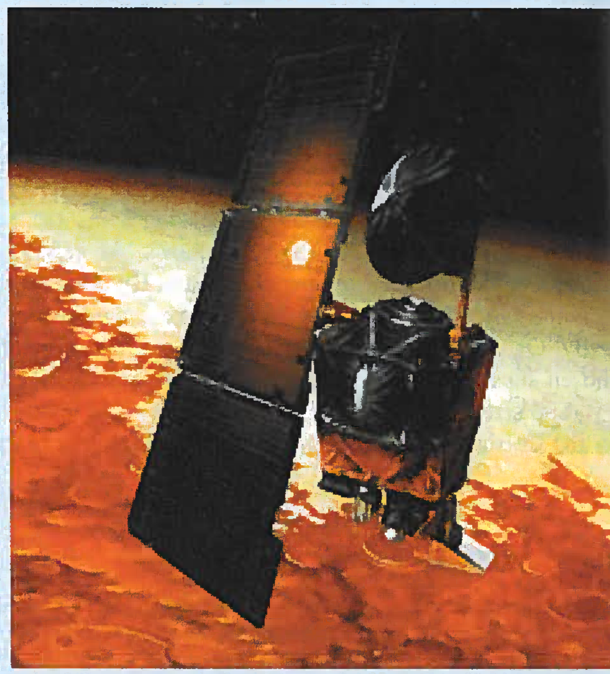
In December 1998 NASA launched the 125-million dollar Mars Climate Orbiter, intended as the red planet's first weather satellite. After a 416-million mi journey, the spacecraft was supposed to go into Mars' orbit on September 23, 1999. Instead, it entered Mars' atmosphere about 100 km (62 mi) lower than planned and was destroyed by heat. The mission controllers said the loss of the spacecraft was due to the failure to convert English measurement units into metric units in the navigation software.

Engineers at Lockheed Martin Corporation who built the spacecraft specified its thrust in pounds, which is an English unit. Scientists at NASA's Jet Propulsion Laboratory, on the other hand, had assumed that thrust data they received were expressed in metric units, as newtons. Normally, pound is the unit for mass. Expressed as a unit for force, however, 1 lb is the force due to gravitational attraction on an object of that mass. To carry out the conversion between pound and newton, we start with 1 lb = 0.4536 kg and from Newton's second law of motion,

$$\begin{aligned} \text{force} &= \text{mass} \times \text{acceleration} \\ &= 0.4536\ \text{kg} \times 9.81\ \text{m/s}^2 \\ &= 4.45\ \text{kg m/s}^2 \\ &= 4.45\ \text{N} \end{aligned}$$

because 1 newton (N) = 1 kg m/s^2. Therefore, instead of converting one pound of force to 4.45 N, the scientists treated it as 1 N.

The considerably smaller engine thrust expressed in newtons resulted in a lower orbit and the ultimate destruction of the spacecraft. Commenting on the failure of the Mars mission, one scientist said: "This is going to be the cautionary tale that will be embedded into introduction to the metric system in elementary school, high school, and college science courses till the end of time."



Artist's conception of the Martian Climate Orbiter.

Both the Celsius and the Kelvin scales have units of equal magnitude; that is, one degree Celsius is equivalent to one kelvin. Experimental studies have shown that absolute zero on the Kelvin scale is equivalent to −273.15°C on the Celsius scale. Thus, we can use the following equation to convert degrees Celsius to kelvin:

$$?\ \text{K} = (^\circ\text{C} + 273.15^\circ\text{C})\,\frac{1\ \text{K}}{1^\circ\text{C}} \qquad (1.4)$$

We will frequently find it necessary to convert between degrees Celsius and degrees Fahrenheit and between degrees Celsius and kelvin. Example 1.3 illustrates these conversions.

The Chemistry in Action essay above shows why we must be careful with units in scientific work.

Example 1.3

(a) Solder is an alloy made of tin and lead that is used in electronic circuits. A certain solder has a melting point of 224°C. What is its melting point in degrees Fahrenheit? (b) Helium has the lowest boiling point of all the elements at −452°F. Convert this temperature to degrees Celsius. (c) Mercury, the only metal that exists as a liquid at room temperature, melts at −38.9°C. Convert its melting point to kelvins.

Solution These three parts require that we carry out temperature conversions, so we need Equations (1.2), (1.3), and (1.4). Keep in mind that the lowest temperature on the Kelvin scale is zero (0 K); therefore, it can never be negative.

(a) This conversion is carried out by writing

$$\frac{9°\text{F}}{5°\text{C}} \times (224°\text{C}) + 32°\text{F} = \boxed{435°\text{F}}$$

(b) Here we have

$$(-452°\text{F} - 32°\text{F}) \times \frac{5°\text{C}}{9°\text{F}} = \boxed{-269°\text{C}}$$

(c) The melting point of mercury in kelvins is given by

$$(-38.9°\text{C} + 273.15°\text{C}) \times \frac{1\text{ K}}{1°\text{C}} = \boxed{234.3\text{ K}}$$

Practice Exercise Convert (a) 327.5°C (the melting point of lead) to degrees Fahrenheit; (b) 172.9°F (the boiling point of ethanol) to degrees Celsius; and (c) 77 K, the boiling point of liquid nitrogen, to degrees Celsius.

Solder is used extensively in the construction of electronic circuits.

Similar problems: 1.24, 1.25, 1.26.

1.8 Handling Numbers

Having surveyed some of the units used in chemistry, we now turn to techniques for handling numbers associated with measurements: scientific notation and significant figures.

Scientific Notation

Chemists often deal with numbers that are either extremely large or extremely small. For example, in 1 g of the element hydrogen there are roughly

$$602{,}200{,}000{,}000{,}000{,}000{,}000{,}000$$

hydrogen atoms. Each hydrogen atom has a mass of only

$$0.00000000000000000000000166\text{ g}$$

These numbers are cumbersome to handle, and it is easy to make mistakes when using them in arithmetic computations. Consider the following multiplication:

$$0.0000000056 \times 0.00000000048 = 0.000000000000000002688$$

It would be easy for us to miss one zero or add one more zero after the decimal point. Consequently, when working with very large and very small numbers, we use a system called *scientific notation.* Regardless of their magnitude, all numbers can be expressed in the form

$$N \times 10^n$$

where N is a number between 1 and 10 and n, the exponent, is a positive or negative integer (whole number). Any number expressed in this way is said to be written in scientific notation.

Suppose that we are given a certain number and asked to express it in scientific notation. Basically, this assignment calls for us to find n. We count the number of places that the decimal point must be moved to give the number N (which is between 1 and 10). If the decimal point has to be moved to the left, then n is a positive integer; if it has to be moved to the right, n is a negative integer. The following examples illustrate the use of scientific notation:

(1) Express 568.762 in scientific notation:

$$568.762 = 5.68762 \times 10^2$$

Note that the decimal point is moved to the left by two places and $n = 2$.

(2) Express 0.00000772 in scientific notation:

$$0.00000772 = 7.72 \times 10^{-6}$$

Here the decimal point is moved to the right by six places and $n = -6$.

Any number raised to the power zero is equal to one.

Keep in mind the following two points. First, $n = 0$ is used for numbers that are not expressed in scientific notation. For example, 74.6×10^0 ($n = 0$) is equivalent to 74.6. Second, the usual practice is to omit the superscript when $n = 1$. Thus the scientific notation for 74.6 is 7.46×10 and not 7.46×10^1.

Next, we consider how scientific notation is handled in arithmetic operations.

Addition and Subtraction

To add or subtract using scientific notation, we first write each quantity—say N_1 and N_2—with the same exponent n. Then we combine N_1 and N_2; the exponents remain the same. Consider the following examples:

$$\begin{aligned}
(7.4 \times 10^3) + (2.1 \times 10^3) &= 9.5 \times 10^3 \\
(4.31 \times 10^4) + (3.9 \times 10^3) &= (4.31 \times 10^4) + (0.39 \times 10^4) \\
&= 4.70 \times 10^4 \\
(2.22 \times 10^{-2}) - (4.10 \times 10^{-3}) &= (2.22 \times 10^{-2}) - (0.41 \times 10^{-2}) \\
&= 1.81 \times 10^{-2}
\end{aligned}$$

Multiplication and Division

To multiply numbers expressed in scientific notation, we multiply N_1 and N_2 in the usual way, but *add* the exponents together. To divide using scientific notation, we divide N_1 and N_2 as usual and subtract the exponents. The following examples show how these operations are performed:

$$\begin{aligned}
(8.0 \times 10^4) \times (5.0 \times 10^2) &= (8.0 \times 5.0)(10^{4+2}) \\
&= 40 \times 10^6 \\
&= 4.0 \times 10^7
\end{aligned}$$

$$(4.0 \times 10^{-5}) \times (7.0 \times 10^{3}) = (4.0 \times 7.0)(10^{-5+3})$$
$$= 28 \times 10^{-2}$$
$$= 2.8 \times 10^{-1}$$

$$\frac{6.9 \times 10^{7}}{3.0 \times 10^{-5}} = \frac{6.9}{3.0} \times 10^{7-(-5)}$$
$$= 2.3 \times 10^{12}$$

$$\frac{8.5 \times 10^{4}}{5.0 \times 10^{9}} = \frac{8.5}{5.0} \times 10^{4-9}$$
$$= 1.7 \times 10^{-5}$$

Significant Figures

Except when all the numbers involved are integers (for example, in counting the number of students in a class), it is often impossible to obtain the exact value of the quantity under investigation. For this reason, it is important to indicate the margin of error in a measurement by clearly indicating the number of ***significant figures,*** which are *the meaningful digits in a measured or calculated quantity.* When significant figures are used, the last digit is understood to be uncertain. For example, we might measure the volume of a given amount of liquid using a graduated cylinder with a scale that gives an uncertainty of 1 mL in the measurement. If the volume is found to be 6 mL, then the actual volume is in the range of 5 mL to 7 mL. We represent the volume of the liquid as (6 ± 1) mL. In this case, there is only one significant figure (the digit 6) that is uncertain by either plus or minus 1 mL. For greater accuracy, we might use a graduated cylinder that has finer divisions, so that the volume we measure is now uncertain by only 0.1 mL. If the volume of the liquid is now found to be 6.0 mL, we may express the quantity as (6.0 ± 0.1) mL, and the actual value is somewhere between 5.9 mL and 6.1 mL. We can further improve the measuring device and obtain more significant figures, but in every case, the last digit is always uncertain; the amount of this uncertainty depends on the particular measuring device we use.

Figure 1.12 shows a modern balance. Balances such as this one are available in many general chemistry laboratories; they readily measure the mass of objects to four decimal places. Therefore, the measured mass typically will have four significant figures (for example, 0.8642 g) or more (for example, 3.9745 g). Keeping track of the number of significant figures in a measurement such as mass ensures that calculations involving the data will reflect the precision of the measurement.

Guidelines for Using Significant Figures

We must always be careful in scientific work to write the proper number of significant figures. In general, it is fairly easy to determine how many significant figures a number has by following these rules:

1. Any digit that is not zero is significant. Thus, 845 cm has three significant figures, 1.234 kg has four significant figures, and so on.
2. Zeros between nonzero digits are significant. Thus, 606 m contains three significant figures, 40,501 kg contains five significant figures, and so on.
3. Zeros to the left of the first nonzero digit are not significant. Their purpose is to indicate the placement of the decimal point. For example, 0.08 L contains one significant figure, 0.0000349 g contains three significant figures, and so on.
4. If a number is greater than 1, then all the zeros written to the right of the decimal point count as significant figures. Thus, 2.0 mg has two significant figures

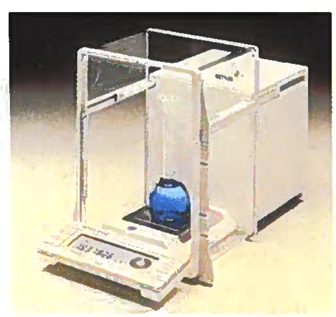


Figure 1.12 *A single-pan balance.*

40.062 mL has five significant figures, and 3.040 dm has four significant figures. If a number is less than 1, then only the zeros that are at the end of the number and the zeros that are between nonzero digits are significant. This means that 0.090 kg has two significant figures, 0.3005 L has four significant figures, 0.00420 min has three significant figures, and so on.

5. For numbers that do not contain decimal points, the trailing zeros (that is, zeros after the last nonzero digit) may or may not be significant. Thus, 400 cm may have one significant figure (the digit 4), two significant figures (40), or three significant figures (400). We cannot know which is correct without more information. By using scientific notation, however, we avoid this ambiguity. In this particular case, we can express the number 400 as 4×10^2 for one significant figure, 4.0×10^2 for two significant figures, or 4.00×10^2 for three significant figures.

Example 1.4 shows the determination of significant figures.

Example 1.4

Determine the number of significant figures in the following measurements: (a) 478 cm, (b) 6.01 g, (c) 0.825 m, (d) 0.043 kg, (e) 1.310×10^{22} atoms, (f) 7000 mL.

Solution (a) Three, because each digit is a nonzero digit. (b) Three, because zeros between nonzero digits are significant. (c) Three, because zeros to the left of the first nonzero digit do not count as significant figures. (d) Two. Same reason as in (c). (e) Four, because the number is greater than one so all the zeros written to the right of the decimal point count as significant figures. (f) This is an ambiguous case. The number of significant figures may be four (7.000×10^3), three (7.00×10^3), two (7.0×10^3), or one (7×10^3). This example illustrates why scientific notation must be used to show the proper number of significant figures.

Similar problems: 1.33, 1.34.

Practice Exercise Determine the number of significant figures in each of the following measurements: (a) 24 mL, (b) 3001 g, (c) 0.0320 m^3, (d) 6.4×10^4 molecules, (e) 560 kg.

A second set of rules specifies how to handle significant figures in calculations.

1. In addition and subtraction, the answer cannot have more digits to the right of the decimal point than either of the original numbers. Consider these examples:

$$\begin{array}{rl} 89.332 & \\ +\ 1.1 & \longleftarrow \text{one digit after the decimal point} \\ \hline 90.432 & \longleftarrow \text{round off to 90.4} \end{array}$$

$$\begin{array}{rl} 2.097 & \\ -0.12 & \longleftarrow \text{two digits after the decimal point} \\ \hline 1.977 & \longleftarrow \text{round off to 1.98} \end{array}$$

The rounding-off procedure is as follows. To round off a number at a certain point we simply drop the digits that follow if the first of them is less than 5. Thus, 8.724 rounds off to 8.72 if we want only two digits after the decimal point. If the first digit following the point of rounding off is equal to or greater than 5, we add 1 to the preceding digit. Thus, 8.727 rounds off to 8.73, and 0.425 rounds off to 0.43.

2. In multiplication and division, the number of significant figures in the final product or quotient is determined by the original number that has the *smallest* number of significant figures. The following examples illustrate this rule:

$$2.8 \times 4.5039 = 12.61092 \longleftarrow \text{round off to } 13$$

$$\frac{6.85}{112.04} = 0.0611388789 \longleftarrow \text{round off to } 0.0611$$

3. Keep in mind that *exact numbers* obtained from definitions or by counting numbers of objects can be considered to have an infinite number of significant figures. For example, the inch is defined to be exactly 2.54 centimeters; that is,

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

Thus, the "2.54" in the equation should not be interpreted as a measured number with three significant figures. In calculations involving conversion between in and cm, we treat both "1" and "2.54" as having an infinite number of significant figures. Similarly, if an object has a mass of 5.0 g, then the mass of nine such objects is

$$5.0 \text{ g} \times 9 = 45 \text{ g}$$

The answer has two significant figures because 5.0 g has two significant figures. The number 9 is exact and does not determine the number of significant figures.

Example 1.5 shows how significant figures are handled in arithmetic operations.

Example 1.5

Carry out the following arithmetic operations to the correct number of significant figures: (a) 11,254.1 g + 0.1983 g, (b) 66.59 L − 3.113 L, (c) 8.16 m × 5.1355, (d) 0.0154 kg ÷ 88.3 mL, (e) 2.64×10^3 cm + 3.27×10^2 cm.

Solution In addition and subtraction, the number of decimal places in the answer is determined by the number having the lowest number of decimal places. In multiplication and division, the significant number of the answer is determined by the number having the smallest number of significant figures.

(a)
$$\begin{array}{r} 11{,}254.1 \text{ g} \\ +\quad 0.1983 \text{ g} \\ \hline 11{,}254.2983 \text{ g} \end{array} \longleftarrow \text{round off to } 11{,}254.3 \text{ g}$$

(b)
$$\begin{array}{r} 66.59 \text{ L} \\ -\ 3.113 \text{ L} \\ \hline 63.477 \text{ L} \end{array} \longleftarrow \text{round off to } 63.48 \text{ L}$$

(c) $8.16 \text{ m} \times 5.1355 = 41.90568 \text{ m} \longleftarrow \text{round off to } 41.9 \text{ m}$

(d) $\dfrac{0.0154 \text{ kg}}{88.3 \text{ mL}} = 0.000174405436 \text{ kg/mL} \longleftarrow \text{round off to } 0.000174 \text{ kg/mL}$ or 1.74×10^{-4} kg/mL

(e) First we change 3.27×10^2 cm to 0.327×10^3 cm and then carry out the addition $(2.64 \text{ cm} + 0.327 \text{ cm}) \times 10^3$. Following the procedure in (a), we find the answer is 2.97×10^3 cm.

Similar problems: 1.35, 1.36.

Practice Exercise Carry out the following arithmetic operations and round off the answers to the appropriate number of significant figures: (a) 26.5862 L + 0.17 L, (b) 9.1 g − 4.682 g, (c) 7.1×10^4 dm × 2.2654×10^2 dm, (d) 6.54 g ÷ 86.5542 mL, (e) $(7.55 \times 10^4 \text{ m}) - (8.62 \times 10^3 \text{ m})$.

The preceding rounding-off procedure applies to one-step calculations. In *chain calculations,* that is, calculations involving more than one step, we can get a different answer depending on how we round off. Consider the following two-step calculations:

First step: $A \times B = C$
Second step: $C \times D = E$

Let's suppose that $A = 3.66$, $B = 8.45$, and $D = 2.11$. Depending on whether we round off C to three or four significant figures, we obtain a different number for E:

Method 1	Method 2
$3.66 \times 8.45 = 30.9$	$3.66 \times 8.45 = 30.93$
$30.9 \times 2.11 = 65.2$	$30.93 \times 2.11 = 65.3$

However, if we had carried out the calculation as $3.66 \times 8.45 \times 2.11$ on a calculator without rounding off the intermediate answer, we would have obtained 65.3 as the answer for E. Although retaining an additional digit past the number of significant figures for intermediate steps helps to eliminate errors from rounding, this procedure is not necessary for most calculations because the difference between the answers is usually quite small. Therefore, for most examples and end-of-chapter problems where intermediate answers are reported, all answers, intermediate and final, will be rounded.

Accuracy and Precision

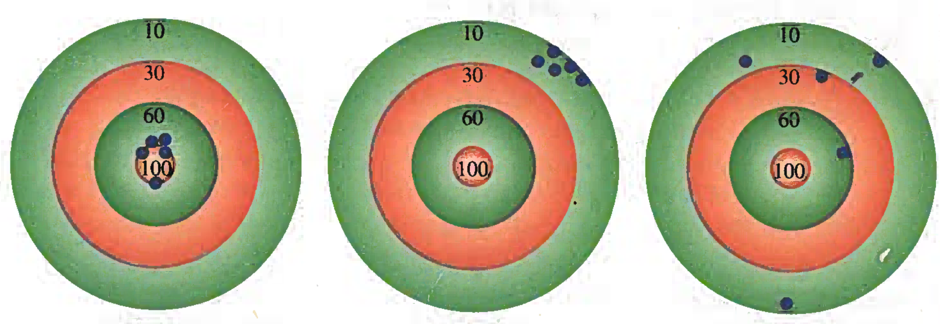
Interactivity:
Accuracy and Precision
ARIS, Interactives

In discussing measurements and significant figures it is useful to distinguish between *accuracy* and *precision.* ***Accuracy*** tells us *how close a measurement is to the true value of the quantity that was measured.* To a scientist there is a distinction between accuracy and precision. ***Precision*** *refers to how closely two or more measurements of the same quantity agree with one another* (Figure 1.13).

The difference between accuracy and precision is a subtle but important one. Suppose, for example, that three students are asked to determine the mass of a piece of copper wire. The results of two successive weighings by each student are

	Student A	Student B	Student C
	1.964 g	1.972 g	2.000 g
	1.978 g	1.968 g	2.002 g
Average value	1.971 g	1.970 g	2.001 g

The true mass of the wire is 2.000 g. Therefore, Student B's results are more *precise* than those of Student A (1.972 g and 1.968 g deviate less from 1.970 g than 1.964 g



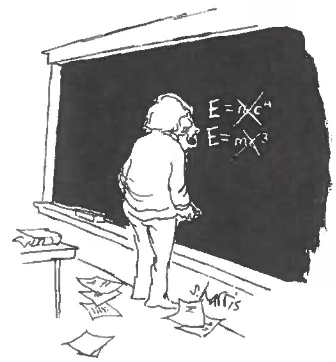
(a) (b) (c)

Figure 1.13 *The distribution of darts on a dart board shows the difference between precise and accurate. (a) Good accuracy and good precision. (b) Poor accuracy and good precision. (c) Poor accuracy and poor precision. The blue dots show the positions of the darts.*

and 1.978 g from 1.971 g), but neither set of results is very *accurate.* Student C's results are not only the most *precise,* but also the most *accurate,* because the average value is closest to the true value. Highly accurate measurements are usually precise too. On the other hand, highly precise measurements do not necessarily guarantee accurate results. For example, an improperly calibrated meterstick or a faulty balance may give precise readings that are in error.

1.9 Dimensional Analysis in Solving Problems

Careful measurements and the proper use of significant figures, along with correct calculations, will yield accurate numerical results. But to be meaningful, the answers also must be expressed in the desired units. The procedure we use to convert between units in solving chemistry problems is called *dimensional analysis* (also called the *factor-label method*). A simple technique requiring little memorization, dimensional analysis is based on the relationship between different units that express the same physical quantity. For example, we know that the monetary unit "dollar" is different from the unit "penny." However, 1 dollar is *equivalent* to 100 pennies because they both represent the same amount of money; that is,



Dimensional analysis might also have led Einstein to his famous mass-energy equation ($E = mc^2$).

Interactivity: Dimensional Analysis ARIS, Interactives

$$1 \text{ dollar} = 100 \text{ pennies}$$

This equivalence enables us to write a conversion factor

$$\frac{1 \text{ dollar}}{100 \text{ pennies}}$$

if we want to convert pennies to dollars. Conversely, the conversion factor

$$\frac{100 \text{ pennies}}{1 \text{ dollar}}$$

enables us to convert dollars to pennies. A conversion factor, then, is a fraction whose numerator and denominator are the same quantity expressed in different units.

Now consider the problem

$$? \text{ pennies} = 2.46 \text{ dollars}$$

Because this is a dollar-to-penny conversion, we choose the conversion factor that has the unit "dollar" in the denominator (to cancel the "dollars" in 2.46 dollars) and write

$$2.46 \cancel{\text{dollars}} \times \frac{100 \text{ pennies}}{1 \cancel{\text{dollar}}} = 246 \text{ pennies}$$

Note that the conversion factor 100 pennies/1 dollar contains exact numbers, so it does not affect the number of significant figures in the final answer.

Next let us consider the conversion of 57.8 meters to centimeters. This problem can be expressed as

$$? \text{ cm} = 57.8 \text{ m}$$

By definition,

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

Because we are converting "m" to "cm," we choose the conversion factor that has meters in the denominator,

$$\frac{1\ \text{cm}}{1 \times 10^{-2}\ \text{m}}$$

and write the conversion as

$$\begin{aligned} ?\ \text{cm} &= 57.8\ \cancel{\text{m}} \times \frac{1\ \text{cm}}{1 \times 10^{-2}\ \cancel{\text{m}}} \\ &= 5780\ \text{cm} \\ &= 5.78 \times 10^{3}\ \text{cm} \end{aligned}$$

Note that scientific notation is used to indicate that the answer has three significant figures. Again, the conversion factor 1 cm/1 $\times$ 10^{-2} m contains exact numbers; therefore, it does not affect the number of significant figures.

In general, to apply dimensional analysis we use the relationship

$$\text{given quantity} \times \text{conversion factor} = \text{desired quantity}$$

and the units cancel as follows:

Remember that the unit we want appears in the numerator and the unit we want to cancel appears in the denominator.

$$\cancel{\text{given unit}} \times \frac{\text{desired unit}}{\cancel{\text{given unit}}} = \text{desired unit}$$

In dimensional analysis, the units are carried through the entire sequence of calculations. Therefore, if the equation is set up correctly, then all the units will cancel except the desired one. If this is not the case, then an error must have been made somewhere, and it can usually be spotted by reviewing the solution.

A Note on Problem Solving

At this point you have been introduced to scientific notation, significant figures, and dimensional analysis, which will help you in solving numerical problems. Chemistry is an experimental science and many of the problems are quantitative in nature. The key to success in problem solving is practice. Just as a marathon runner cannot prepare for a race by simply reading books on running and a pianist cannot give a successful concert by only memorizing the musical score, you cannot be sure of your understanding of chemistry without solving problems. The following steps will help to improve your skill at solving numerical problems.

1. Read the question carefully. Understand the information that is given and what you are asked to solve. Frequently it is helpful to make a sketch that will help you to visualize the situation.
2. Find the appropriate equation that relates the given information and the unknown quantity. Sometimes solving a problem will involve more than one step, and you may be expected to look up quantities in tables that are not provided in the problem. Dimensional analysis is often needed to carry out conversions.
3. Check your answer for the correct sign, units, and significant figures.
4. A very important part of problem solving is being able to judge whether the answer is reasonable. It is relatively easy to spot a wrong sign or incorrect units. But if a number (say 9) is incorrectly placed in the denominator instead of in the numerator, the answer would be too small even if the sign and units of the calculated quantity were correct.

5. One way to quickly check the answer is to make a "ball-park" estimate. The idea here is to round off the numbers in the calculation in such a way so as to simplify the arithmetic. This approach is sometimes called the "back-of-the-envelope calculation" because it can be done easily without using a calculator. The answer you get will not be exact, but it will be close to the correct one.

Example 1.6

A person's average daily intake of glucose (a form of sugar) is 0.0833 pound (lb). What is this mass in milligrams (mg)? (1 lb = 453.6 g.)

Conversion factors for some of the English system units commonly used in the United States for nonscientific measurements (for example, pounds and inches) are provided inside the back cover of this book.

Strategy The problem can be stated as

$$? \text{ mg} = 0.0833 \text{ lb}$$

The relationship between pounds and grams is given in the problem. This relationship will enable conversion from pounds to grams. A metric conversion is then needed to convert grams to milligrams (1 mg = 1×10^{-3} g). Arrange the appropriate conversion factors so that pounds and grams cancel and the unit milligrams is obtained in your answer.

Solution The sequence of conversions is

$$\text{pounds} \longrightarrow \text{grams} \longrightarrow \text{milligrams}$$

Using the following conversion factors

$$\frac{453.6 \text{ g}}{1 \text{ lb}} \quad \text{and} \quad \frac{1 \text{ mg}}{1 \times 10^{-3} \text{ g}}$$

we obtain the answer in one step:

$$? \text{ mg} = 0.0833 \cancel{\text{lb}} \times \frac{453.6 \cancel{\text{g}}}{1 \cancel{\text{lb}}} \times \frac{1 \text{ mg}}{1 \times 10^{-3} \cancel{\text{g}}} = \boxed{3.78 \times 10^4 \text{ mg}}$$

Check As an estimate, we note that 1 lb is roughly 500 g and that 1 g = 1000 mg. Therefore, 1 lb is roughly 5×10^5 mg. Rounding off 0.0833 lb to 0.1 lb, we get 5×10^4 mg, which is close to the preceding quantity.

Similar problem: 1.43.

Practice Exercise A roll of aluminum foil has a mass of 1.07 kg. What is its mass in pounds?

As Examples 1.7 and 1.8 illustrate, conversion factors can be squared or cubed in dimensional analysis.

Example 1.7

An average adult has 5.2 L of blood. What is the volume of blood in m^3?

Strategy The problem can be stated as

$$? \text{ m}^3 = 5.2 \text{ L}$$

(Continued)

How many conversion factors are needed for this problem? Recall that 1 L = 1000 cm^3 and 1 cm = 1×10^{-2} m.

Solution We need two conversion factors here: one to convert liters to cm^3 and one to convert centimeters to meters:

$$\frac{1000\ cm^3}{1\ L} \quad \text{and} \quad \frac{1 \times 10^{-2}\ m}{1\ cm}$$

Because the second conversion factor deals with length (cm and m) and we want volume here, it must therefore be cubed to give

$$\frac{1 \times 10^{-2}\ m}{1\ cm} \times \frac{1 \times 10^{-2}\ m}{1\ cm} \times \frac{1 \times 10^{-2}\ m}{1\ cm} = \left(\frac{1 \times 10^{-2}\ m}{1\ cm}\right)^3$$

This means that 1 cm^3 = $1 \times 10^{-6}\ m^3$. Now we can write

$$?\ m^3 = 5.2\ \cancel{L} \times \frac{1000\ cm^3}{1\ \cancel{L}} \times \left(\frac{1 \times 10^{-2}\ m}{1\ cm}\right)^3 = \boxed{5.2 \times 10^{-3}\ m^3}$$

Check From the preceding conversion factors you can show that 1 L = $1 \times 10^{-3}\ m^3$. Therefore, 5 L of blood would be equal to $5 \times 10^{-3}\ m^3$, which is close to the answer.

Similar problem: 1.48(d).

Practice Exercise The volume of a room is $1.08 \times 10^8\ dm^3$. What is the volume in m^3?

Liquid nitrogen

Example 1.8

Liquid nitrogen is obtained from liquefied air and is used to prepare frozen goods and in low-temperature research. The density of the liquid at its boiling point (−196°C or 77 K) is 0.808 g/cm^3. Convert the density to units of kg/m^3.

Strategy The problem can be stated as

$$?\ kg/m^3 = 0.808\ g/cm^3$$

Two separate conversions are required for this problem: g $\longrightarrow$ kg and $cm^3 \longrightarrow m^3$. Recall that 1 kg = 1000 g and 1 cm = 1×10^{-2} m.

Solution In Example 1.7 we saw that 1 cm^3 = $1 \times 10^{-6}\ m^3$. The conversion factors are

$$\frac{1\ kg}{1000\ g} \quad \text{and} \quad \frac{1\ cm^3}{1 \times 10^{-6}\ m^3}$$

Finally,

$$?\ kg/m^3 = \frac{0.808\ \cancel{g}}{1\ \cancel{cm^3}} \times \frac{1\ kg}{1000\ \cancel{g}} \times \frac{1\ \cancel{cm^3}}{1 \times 10^{-6}\ m^3} = \boxed{808\ kg/m^3}$$

Check Because 1 m^3 = $1 \times 10^6\ cm^3$, we would expect much more mass in 1 m^3 than in 1 cm^3. Therefore, the answer is reasonable.

Similar problem: 1.49.

Practice Exercise The density of the lightest metal, lithium (Li), is 5.34×10^2 kg/m^3. Convert the density to g/cm^3.

Summary of Facts and Concepts

1. The study of chemistry involves three basic steps: observation, representation, and interpretation. Observation refers to measurements in the macroscopic world; representation involves the use of shorthand notation symbols and equations for communication; interpretations are based on atoms and molecules, which belong to the microscopic world.
2. The scientific method is a systematic approach to research that begins with the gathering of information through observation and measurements. In the process, hypotheses, laws, and theories are devised and tested.
3. Chemists study matter and the changes it undergoes. The substances that make up matter have unique physical properties that can be observed without changing their identity and unique chemical properties that, when they are demonstrated, do change the identity of the substances. Mixtures, whether homogeneous or heterogeneous, can be separated into pure components by physical means.
4. The simplest substances in chemistry are elements. Compounds are formed by the chemical combination of atoms of different elements in fixed proportions.
5. All substances, in principle, can exist in three states: solid, liquid, and gas. The interconversion between these states can be effected by changing the temperature.
6. SI units are used to express physical quantities in all sciences, including chemistry.
7. Numbers expressed in scientific notation have the form $N \times 10^n$, where N is between 1 and 10, and n is a positive or negative integer. Scientific notation helps us handle very large and very small quantities.

Key Words

Accuracy, p. 26
Chemical property, p. 14
Chemistry, p. 4
Compound, p. 12
Density, p. 15
Element, p. 12
Extensive property, p. 15
Heterogeneous mixture, p. 11
Homogeneous mixture, p. 11
Hypothesis, p. 9
Intensive property, p. 15
International System of Units (SI), p. 15
Kelvin, p. 19
Law, p. 9
Liter, p. 17
Macroscopic property, p. 15
Mass, p. 15
Matter, p. 11
Microscopic property, p. 15
Mixture, p. 11
Physical property, p. 14
Precision, p. 26
Qualitative, p. 8
Quantitative, p. 8
Scientific method, p. 8
Significant figures, p. 23
Substance, p. 11
Theory, p. 9
Volume, p. 15
Weight, p. 16

Questions and Problems

The Scientific Method

Review Questions

1.1 Explain what is meant by the scientific method.

1.2 What is the difference between qualitative data and quantitative data?

Problems

1.3 Classify the following as qualitative or quantitative statements, giving your reasons. (a) The sun is approximately 93 million mi from Earth. (b) Leonardo da Vinci was a better painter than Michelangelo. (c) Ice is less dense than water. (d) Butter tastes better than margarine. (e) A stitch in time saves nine.

1.4 Classify each of the following statements as a hypothesis, a law, or a theory. (a) Beethoven's contribution to music would have been much greater if he had married. (b) An autumn leaf gravitates toward the ground because there is an attractive force between the leaf and Earth. (c) All matter is composed of very small particles called atoms.

Classification and Properties of Matter

Review Questions

1.5 Give an example for each of the following terms: (a) matter, (b) substance, (c) mixture.

1.6 Give an example of a homogeneous mixture and an example of a heterogeneous mixture.

1.7 Using examples, explain the difference between a physical property and a chemical property.

1.8 How does an intensive property differ from an extensive property? Which of the following properties are intensive and which are extensive? (a) length, (b) volume, (c) temperature, (d) mass.

1.9 Give an example of an element and a compound. How do elements and compounds differ?

1.10 What is the number of known elements?

Problems

1.11 Do the following statements describe chemical or physical properties? (a) Oxygen gas supports combustion. (b) Fertilizers help to increase agricultural production. (c) Water boils below 100°C on top of a mountain. (d) Lead is denser than aluminum. (e) Uranium is a radioactive element.

1.12 Does each of the following describe a physical change or a chemical change? (a) The helium gas inside a balloon tends to leak out after a few hours. (b) A flashlight beam slowly gets dimmer and finally goes out. (c) Frozen orange juice is reconstituted by adding water to it. (d) The growth of plants depends on the sun's energy in a process called photosynthesis. (e) A spoonful of table salt dissolves in a bowl of soup.

1.13 Give the names of the elements represented by the chemical symbols Li, F, P, Cu, As, Zn, Cl, Pt, Mg, U, Al, Si, Ne. (See Table 1.1 and the inside front cover.)

1.14 Give the chemical symbols for the following elements: (a) potassium, (b) tin, (c) chromium, (d) boron, (e) barium, (f) plutonium, (g) sulfur, (h) argon, (i) mercury. (See Table 1.1 and the inside front cover.)

1.15 Classify each of the following substances as an element or a compound: (a) hydrogen, (b) water, (c) gold, (d) sugar.

1.16 Classify each of the following as an element, a compound, a homogeneous mixture, or a heterogeneous mixture: (a) seawater, (b) helium gas, (c) sodium chloride (table salt), (d) a bottle of soft drink, (e) a milkshake, (f) air in a bottle, (g) concrete.

Measurement

Review Questions

1.17 Name the SI base units that are important in chemistry. Give the SI units for expressing the following: (a) length, (b) volume, (c) mass, (d) time, (e) energy, (f) temperature.

1.18 Write the numbers represented by the following prefixes: (a) mega-, (b) kilo-, (c) deci-, (d) centi-, (e) milli-, (f) micro-, (g) nano-, (h) pico-.

1.19 What units do chemists normally use for density of liquids and solids? For gas density? Explain the differences.

1.20 Describe the three temperature scales used in the laboratory and in everyday life: the Fahrenheit scale, the Celsius scale, and the Kelvin scale.

Problems

1.21 Bromine is a reddish-brown liquid. Calculate its density (in g/mL) if 586 g of the substance occupies 188 mL.

1.22 The density of ethanol, a colorless liquid that is commonly known as grain alcohol, is 0.798 g/mL. Calculate the mass of 17.4 mL of the liquid.

1.23 Convert the following temperatures to degrees Celsius or Fahrenheit: (a) 95°F, the temperature on a hot summer day; (b) 12°F, the temperature on a cold winter day; (c) a 102°F fever; (d) a furnace operating at 1852°F; (e) −273.15°C (theoretically the lowest attainable temperature).

1.24 (a) Normally the human body can endure a temperature of 105°F for only short periods of time without permanent damage to the brain and other vital organs. What is this temperature in degrees Celsius? (b) Ethylene glycol is a liquid organic compound that is used as an antifreeze in car radiators. It freezes at −11.5°C. Calculate its freezing temperature in degrees Fahrenheit. (c) The temperature on the surface of the sun is about 6300°C. What is this temperature in degrees Fahrenheit? (d) The ignition temperature of paper is 451°F. What is the temperature in degrees Celsius?

1.25 Convert the following temperatures to kelvin: (a) 113°C, the melting point of sulfur, (b) 37°C, the normal body temperature, (c) 357°C, the boiling point of mercury.

1.26 Convert the following temperatures to degrees Celsius: (a) 77 K, the boiling point of liquid nitrogen, (b) 4.2 K, the boiling point of liquid helium, (c) 601 K, the melting point of lead.

Handling Numbers

Review Questions

1.27 What is the advantage of using scientific notation over decimal notation?

1.28 Define significant figure. Discuss the importance of using the proper number of significant figures in measurements and calculations.

Problems

1.29 Express the following numbers in scientific notation: (a) 0.000000027, (b) 356, (c) 47,764, (d) 0.096.

1.30 Express the following numbers as decimals: (a) 1.52×10^{-2}, (b) 7.78×10^{-8}.

1.31 Express the answers to the following calculations in scientific notation:

(a) $145.75 + (2.3 \times 10^{-1})$

(b) $79{,}500 \div (2.5 \times 10^{2})$

(c) $(7.0 \times 10^{-3}) - (8.0 \times 10^{-4})$

(d) $(1.0 \times 10^{4}) \times (9.9 \times 10^{6})$

1.32 Express the answers to the following calculations in scientific notation:
(a) $0.0095 + (8.5 \times 10^{-3})$
(b) $653 \div (5.75 \times 10^{-8})$
(c) $850{,}000 - (9.0 \times 10^{5})$
(d) $(3.6 \times 10^{-4}) \times (3.6 \times 10^{6})$

1.33 What is the number of significant figures in each of the following measurements?
(a) 4867 mi
(b) 56 mL
(c) 60,104 ton
(d) 2900 g
(e) 40.2 g/cm^3
(f) 0.0000003 cm
(g) 0.7 min
(h) 4.6×10^{19} atoms

1.34 How many significant figures are there in each of the following? (a) 0.006 L, (b) 0.0605 dm, (c) 60.5 mg, (d) 605.5 cm^2, (e) 960×10^{-3} g, (f) 6 kg, (g) 60 m.

1.35 Carry out the following operations as if they were calculations of experimental results, and express each answer in the correct units with the correct number of significant figures:
(a) 5.6792 m + 0.6 m + 4.33 m
(b) 3.70 g − 2.9133 g
(c) 4.51 cm × 3.6666 cm

1.36 Carry out the following operations as if they were calculations of experimental results, and express each answer in the correct units with the correct number of significant figures:
(a) 7.310 km ÷ 5.70 km
(b) $(3.26 \times 10^{-3}\ \text{mg}) - (7.88 \times 10^{-5}\ \text{mg})$
(c) $(4.02 \times 10^{6}\ \text{dm}) + (7.74 \times 10^{7}\ \text{dm})$

Dimensional Analysis

Problems

1.37 Carry out the following conversions: (a) 22.6 m to decimeters, (b) 25.4 mg to kilograms, (c) 556 mL to liters, (d) 10.6 kg/m^3 to g/cm^3.

1.38 Carry out the following conversions: (a) 242 lb to milligrams, (b) 68.3 cm^3 to cubic meters, (c) 7.2 m^3 to liters, (d) 28.3 μg to pounds.

1.39 The average speed of helium at 25°C is 1255 m/s. Convert this speed to miles per hour (mph).

1.40 How many seconds are there in a solar year (365.24 days)?

1.41 How many minutes does it take light from the sun to reach Earth? (The distance from the sun to Earth is 93 million mi; the speed of light = 3.00×10^{8} m/s.)

1.42 A slow jogger runs a mile in 13 min. Calculate the speed in (a) in/s, (b) m/min, (c) km/h. (1 mi = 1609 m; 1 in = 2.54 cm.)

1.43 A 6.0-ft person weighs 168 lb. Express this person's height in meters and weight in kilograms. (1 lb = 453.6 g; 1 m = 3.28 ft.)

1.44 The current speed limit in some states in the United States is 55 miles per hour. What is the speed limit in kilometers per hour? (1 mi = 1609 m.)

1.45 For a fighter jet to take off from the deck of an aircraft carrier, it must reach a speed of 62 m/s. Calculate the speed in miles per hour (mph).

1.46 The "normal" lead content in human blood is about 0.40 part per million (that is, 0.40 g of lead per million grams of blood). A value of 0.80 part per million (ppm) is considered to be dangerous. How many grams of lead are contained in 6.0×10^{3} g of blood (the amount in an average adult) if the lead content is 0.62 ppm?

1.47 Carry out the following conversions: (a) 1.42 light-years to miles (a light-year is an astronomical measure of distance—the distance traveled by light in a year, or 365 days; the speed of light is 3.00×10^{8} m/s), (b) 32.4 yd to centimeters, (c) 3.0×10^{10} cm/s to ft/s.

1.48 Carry out the following conversions: (a) 185 nm to meters. (b) 4.5 billion years (roughly the age of Earth) to seconds. (Assume there are 365 days in a year.) (c) 71.2 cm^3 to m^3. (d) 88.6 m^3 to liters.

1.49 Aluminum is a lightweight metal (density = 2.70 g/cm^3) used in aircraft construction, high-voltage transmission lines, beverage cans, and foils. What is its density in kg/m^3?

1.50 The density of ammonia gas under certain conditions is 0.625 g/L. Calculate its density in g/cm^3.

Additional Problems

1.51 Give one qualitative and one quantitative statement about each of the following: (a) water, (b) carbon, (c) iron, (d) hydrogen gas, (e) sucrose (cane sugar), (f) table salt (sodium chloride), (g) mercury, (h) gold, (i) air.

1.52 Which of the following statements describe physical properties and which describe chemical properties? (a) Iron has a tendency to rust. (b) Rainwater in industrialized regions tends to be acidic. (c) Hemoglobin molecules have a red color. (d) When a glass of water is left out in the sun, the water gradually disappears. (e) Carbon dioxide in air is converted to more complex molecules by plants during photosynthesis.

1.53 In 2004, about 95.0 billion lb of sulfuric acid were produced in the United States. Convert this quantity to tons.

1.54 In determining the density of a rectangular metal bar, a student made the following measurements: length, 8.53 cm; width, 2.4 cm; height, 1.0 cm; mass, 52.7064 g. Calculate the density of the metal to the correct number of significant figures.

1.55 Calculate the mass of each of the following: (a) a sphere of gold with a radius of 10.0 cm [the volume of a sphere with a radius r is $V = (4/3)\pi r^3$; the density of gold = 19.3 g/cm^3], (b) a cube of platinum of edge length 0.040 mm (the density of platinum = 21.4 g/cm^3), (c) 50.0 mL of ethanol (the density of ethanol = 0.798 g/mL).

1.56 A cylindrical glass tube 12.7 cm in length is filled with mercury. The mass of mercury needed to fill the tube is 105.5 g. Calculate the inner diameter of the tube. (The density of mercury = 13.6 g/mL.)

1.57 The following procedure was used to determine the volume of a flask. The flask was weighed dry and then filled with water. If the masses of the empty flask and filled flask were 56.12 g and 87.39 g, respectively, and the density of water is 0.9976 g/cm^3, calculate the volume of the flask in cm^3.

1.58 The speed of sound in air at room temperature is about 343 m/s. Calculate this speed in miles per hour. (1 mi = 1609 m.)

1.59 A piece of silver (Ag) metal weighing 194.3 g is placed in a graduated cylinder containing 242.0 mL of water. The volume of water now reads 260.5 mL. From these data calculate the density of silver.

1.60 The experiment described in Problem 1.59 is a crude but convenient way to determine the density of some solids. Describe a similar experiment that would enable you to measure the density of ice. Specifically, what would be the requirements for the liquid used in your experiment?

1.61 A lead sphere has a mass of 1.20×10^4 g, and its volume is 1.05×10^3 cm^3. Calculate the density of lead.

1.62 Lithium is the least dense metal known (density: 0.53 g/cm^3). What is the volume occupied by 1.20×10^3 g of lithium?

1.63 The medicinal thermometer commonly used in homes can be read ±0.1°F, whereas those in the doctor's office may be accurate to ±0.1°C. In degrees Celsius, express the percent error expected from each of these thermometers in measuring a person's body temperature of 38.9°C.

1.64 Vanillin (used to flavor vanilla ice cream and other foods) is the substance whose aroma the human nose detects in the smallest amount. The threshold limit is 2.0×10^{-11} g per liter of air. If the current price of 50 g of vanillin is \$112, determine the cost to supply enough vanillin so that the aroma could be detected in a large aircraft hangar with a volume of 5.0×10^7 ft^3.

1.65 At what temperature does the numerical reading on a Celsius thermometer equal that on a Fahrenheit thermometer?

1.66 Suppose that a new temperature scale has been devised on which the melting point of ethanol (−117.3°C) and the boiling point of ethanol (78.3°C) are taken as 0°S and 100°S, respectively, where S is the symbol for the new temperature scale. Derive an equation relating a reading on this scale to a reading on the Celsius scale. What would this thermometer read at 25°C?

1.67 A resting adult requires about 240 mL of pure oxygen/min and breathes about 12 times every minute. If inhaled air contains 20 percent oxygen by volume and exhaled air 16 percent, what is the volume of air per breath? (Assume that the volume of inhaled air is equal to that of exhaled air.)

1.68 (a) Referring to Problem 1.67, calculate the total volume (in liters) of air an adult breathes in a day. (b) In a city with heavy traffic, the air contains 2.1×10^{-6} L of carbon monoxide (a poisonous gas) per liter. Calculate the average daily intake of carbon monoxide in liters by a person.

1.69 The total volume of seawater is 1.5×10^{21} L. Assume that seawater contains 3.1 percent sodium chloride by mass and that its density is 1.03 g/mL. Calculate the total mass of sodium chloride in kilograms and in tons. (1 ton = 2000 lb; 1 lb = 453.6 g.)

1.70 Magnesium (Mg) is a valuable metal used in alloys, in batteries, and in the manufacture of chemicals. It is obtained mostly from seawater, which contains about 1.3 g of Mg for every kilogram of seawater. Referring to Problem 1.69, calculate the volume of seawater (in liters) needed to extract 8.0×10^4 tons of Mg, which is roughly the annual production in the United States.

1.71 A student is given a crucible and asked to prove whether it is made of pure platinum. She first weighs the crucible in air and then weighs it suspended in water (density = 0.9986 g/mL). The readings are 860.2 g and 820.2 g, respectively. Based on these measurements and given that the density of platinum is 21.45 g/cm^3, what should her conclusion be? (*Hint:* An object suspended in a fluid is buoyed up by the mass of the fluid displaced by the object. Neglect the buoyance of air.)

1.72 The surface area and average depth of the Pacific Ocean are 1.8×10^8 km^2 and 3.9×10^3 m, respectively. Calculate the volume of water in the ocean in liters.

1.73 The unit "troy ounce" is often used for precious metals such as gold (Au) and platinum (Pt). (1 troy ounce = 31.103 g.) (a) A gold coin weighs 2.41 troy ounces. Calculate its mass in grams. (b) Is a troy ounce heavier or lighter than an ounce? (1 lb = 16 oz; 1 lb = 453.6 g.)

1.74 Osmium (Os) is the densest element known (density = 22.57 g/cm^3). Calculate the mass in pounds and in kilograms of an Os sphere 15 cm in diameter (about the size of a grapefruit). See Problem 1.55 for volume of a sphere.

1.75 Percent error is often expressed as the absolute value of the difference between the true value and the experimental value, divided by the true value:

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

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The vertical lines indicate absolute value. Calculate the percent error for the following measurements: (a) The density of alcohol (ethanol) is found to be 0.802 g/mL. (True value: 0.798 g/mL.) (b) The mass of gold in an earring is analyzed to be 0.837 g. (True value: 0.864 g.)

1.76 The natural abundances of elements in the human body, expressed as percent by mass, are: oxygen (O), 65 percent; carbon (C), 18 percent; hydrogen (H), 10 percent; nitrogen (N), 3 percent; calcium (Ca), 1.6 percent; phosphorus (P), 1.2 percent; all other elements, 1.2 percent. Calculate the mass in grams of each element in the body of a 62-kg person.

1.77 The men's world record for running a mile outdoors (as of 1997) is 3 min 44.39 s. At this rate, how long would it take to run a 1500-m race? (1 mi = 1609 m.)

1.78 Venus, the second closest planet to the sun, has a surface temperature of 7.3×10^2 K. Convert this temperature to °C and °F.

1.79 Chalcopyrite, the principal ore of copper (Cu), contains 34.63 percent Cu by mass. How many grams of Cu can be obtained from 5.11×10^3 kg of the ore?

1.80 It has been estimated that 8.0×10^4 tons of gold (Au) have been mined. Assume gold costs $350 per ounce. What is the total worth of this quantity of gold?

1.81 A 1.0-mL volume of seawater contains about 4.0×10^{-12} g of gold. The total volume of ocean water is 1.5×10^{21} L. Calculate the total amount of gold (in grams) that is present in seawater, and the worth of the gold in dollars (see Problem 1.80). With so much gold out there, why hasn't someone become rich by mining gold from the ocean?

1.82 Measurements show that 1.0 g of iron (Fe) contains 1.1×10^{22} Fe atoms. How many Fe atoms are in 4.9 g of Fe, which is the total amount of iron in the body of an average adult?

1.83 The thin outer layer of Earth, called the crust, contains only 0.50 percent of Earth's total mass and yet is the source of almost all the elements (the atmosphere provides elements such as oxygen, nitrogen, and a few other gases). Silicon (Si) is the second most abundant element in Earth's crust (27.2 percent by mass). Calculate the mass of silicon in kilograms in Earth's crust. (The mass of Earth is 5.9×10^{21} tons. 1 ton = 2000 lb; 1 lb = 453.6 g.)

1.84 The diameter of a copper (Cu) atom is roughly 1.3×10^{-10} m. How many times can you divide evenly a piece of 10-cm copper wire until it is reduced to two separate copper atoms? (Assume there are appropriate tools for this procedure and that copper atoms are lined up in a straight line, in contact with each other. Round off your answer to an integer.)

1.85 One gallon of gasoline in an automobile's engine produces on the average 9.5 kg of carbon dioxide, which is a greenhouse gas, that is, it promotes the warming of Earth's atmosphere. Calculate the annual production of carbon dioxide in kilograms if there are 40 million cars in the United States and each car covers a distance of 5000 mi at a consumption rate of 20 miles per gallon.

1.86 A sheet of aluminum (Al) foil has a total area of 1.000 ft^2 and a mass of 3.636 g. What is the thickness of the foil in millimeters? (Density of Al = 2.699 g/cm^3.)

1.87 Comment on whether each of the following is a homogeneous mixture or a heterogeneous mixture: (a) air in a closed bottle and (b) air over New York City.

1.88 Chlorine is used to disinfect swimming pools. The accepted concentration for this purpose is 1 ppm chlorine, or 1 g of chlorine per million grams of water. Calculate the volume of a chlorine solution (in milliliters) a homeowner should add to her swimming pool if the solution contains 6.0 percent chlorine by mass and there are 2.0×10^4 gallons of water in the pool. (1 gallon = 3.79 L; density of liquids = 1.0 g/mL.)

1.89 The world's total petroleum reserve is estimated at 2.0×10^{22} J (joule is the unit of energy where 1 J = 1 kg m^2/s^2). At the present rate of consumption, 1.8×10^{20} J/yr, how long would it take to exhaust the supply?

1.90 In water conservation, chemists spread a thin film of certain inert material over the surface of water to cut down the rate of evaporation of water in reservoirs. This technique was pioneered by Benjamin Franklin three centuries ago. Franklin found that 0.10 mL of oil could spread over the surface of water of about 40 m^2 in area. Assuming that the oil forms a *monolayer,* that is, a layer that is only one molecule thick, estimate the length of each oil molecule in nanometers. (1 nm = 1×10^{-9} m.)

1.91 Fluoridation is the process of adding fluorine compounds to drinking water to help fight tooth decay. A concentration of 1 ppm of fluorine is sufficient for the purpose. (1 ppm means one part per million, or 1 g of fluorine per 1 million g of water.) The compound normally chosen for fluoridation is sodium fluoride, which is also added to some toothpastes. Calculate the quantity of sodium fluoride in kilograms needed per year for a city of 50,000 people if the daily consumption of water per person is 150 gallons. What percent of the sodium fluoride is "wasted" if each person uses only 6.0 L of water a day for drinking and cooking? (Sodium fluoride is 45.0 percent fluorine by mass. 1 gallon = 3.79 L; 1 year = 365 days; density of water = 1.0 g/mL.)

1.92 A gas company in Massachusetts charges $1.30 for 15.0 ft^3 of natural gas. (a) Convert this rate to dollars per liter of gas. (b) If it takes 0.304 ft^3 of gas to boil a liter of water, starting at room temperature (25°C), how much would it cost to boil a 2.1-L kettle of water?

1.93 Pheromones are compounds secreted by females of many insect species to attract mates. Typically, 1.0×10^{-8} g of a pheromone is sufficient to reach all targeted males within a radius of 0.50 mi. Calculate the density of the pheromone (in grams per liter) in a circular air space having a radius of 0.50 mi and a height of 40 ft.

Special Problems

1.94 A bank teller is asked to assemble "one-dollar" sets of coins for his clients. Each set is made of three quarters, one nickel, and two dimes. The masses of the coins are: quarter: 5.645 g; nickel: 4.967 g; dime: 2.316 g. What is the maximum number of sets that can be assembled from 33.871 kg of quarters, 10.432 kg of nickels, and 7.990 kg of dimes? What is the total mass (in g) of this collection of coins?

1.95 A graduated cylinder is filled to the 40.00-mL mark with a mineral oil. The masses of the cylinder before and after the addition of the mineral oil are 124.966 g and 159.446 g, respectively. In a separate experiment, a metal ball bearing of mass 18.713 g is placed in the cylinder and the cylinder is again filled to the 40.00-mL mark with the mineral oil. The combined mass of the ball bearing and mineral oil is 50.952 g. Calculate the density and radius of the ball bearing. [The volume of a sphere of radius r is $(4/3)\pi r^3$.]

1.96 Bronze is an alloy made of copper (Cu) and tin (Sn). Calculate the mass of a bronze cylinder of radius 6.44 cm and length 44.37 cm. The composition of the bronze is 79.42 percent Cu and 20.58 percent Sn and the densities of Cu and Sn are 8.94 g/cm^3 and 7.31 g/cm^3, respectively. What assumption should you make in this calculation?

1.97 A chemist in the nineteenth century prepared an unknown substance. In general, do you think it would be more difficult to prove that it is an element or a compound? Explain.

1.98 A chemist mixes two liquids A and B to form a homogeneous mixture. The densities of the liquids are 2.0514 g/mL for A and 2.6678 g/mL for B. When she drops a small object into the mixture, she finds that the object becomes suspended in the liquid; that is, it neither sinks nor floats. If the mixture is made of 41.37 percent A and 58.63 percent B by volume, what is the density of the metal? Can this procedure be used in general to determine the densities of solids? What assumptions must be made in applying this method?

1.99 You are given a liquid. Briefly describe steps you would take to show whether it is a pure substance or a homogeneous mixture.

1.100 Tums is a popular remedy for acid indigestion. A typical Tums tablet contains calcium carbonate plus some inert substances. When ingested, it reacts with the gastric juice (hydrochloric acid) in the stomach to give off carbon dioxide gas. When a 1.328-g tablet reacted with 40.00 mL of hydrochloric acid (density: 1.140 g/mL), carbon dioxide gas was given off and the resulting solution weighed 46.699 g. Calculate the number of liters of carbon dioxide gas released if its density is 1.81 g/L.

1.101 A 250-mL glass bottle was filled with 242 mL of water at 20°C and tightly capped. It was then left outdoors overnight, where the average temperature was −5°C. Predict what would happen. The density of water at 20°C is 0.998 g/cm^3 and that of ice at −5°C is 0.916 g/cm^3.

Answers to Practice Exercises

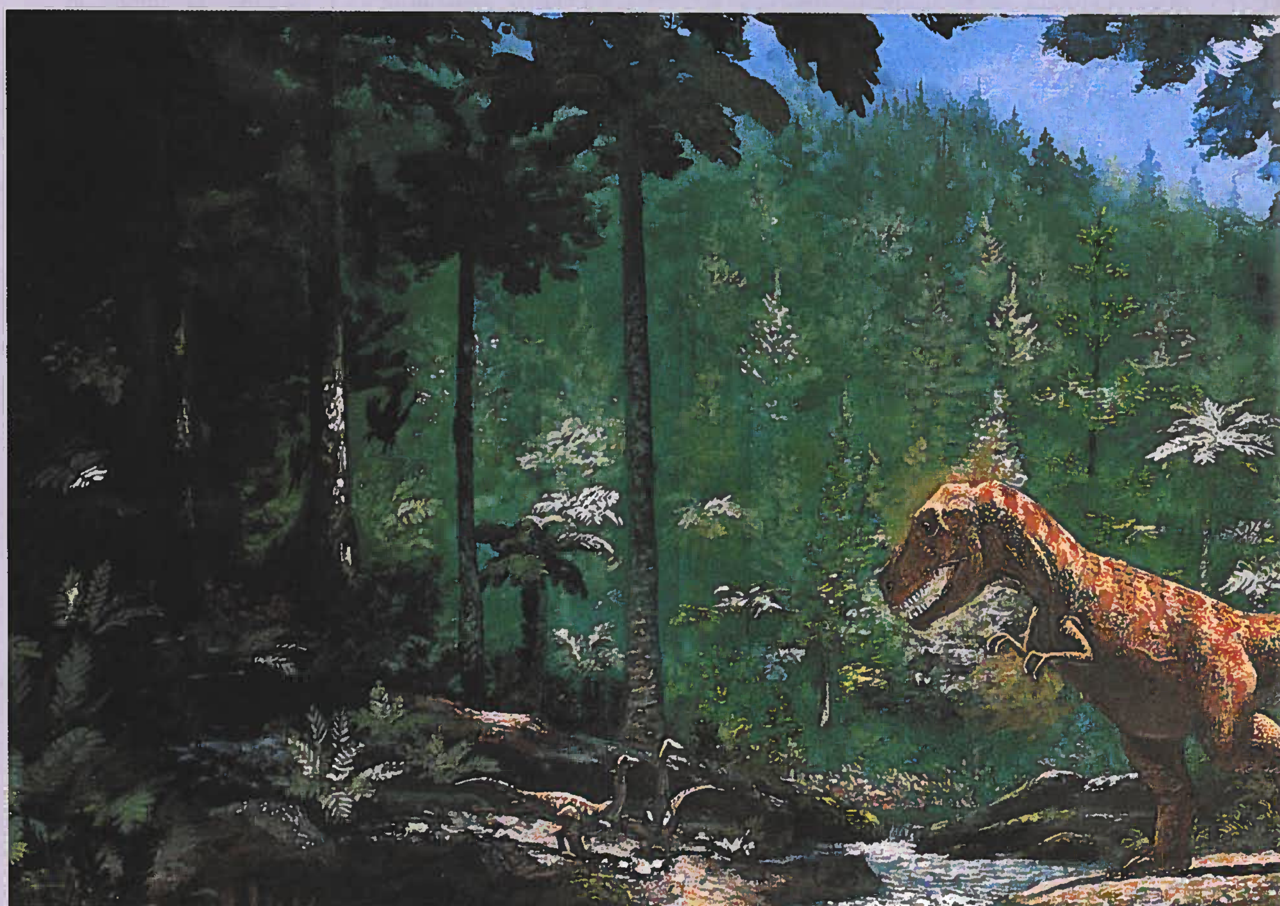
1.1 96.5 g. **1.2** 341 g. **1.3** (a) 621.5°F, (b) 78.3°C, (c) −196°C. **1.4** (a) Two, (b) four, (c) three, (d) two, (e) three or two. **1.5** (a) 26.76 L, (b) 4.4 g, (c) 1.6×10^7 dm^2, (d) 0.0756 g/mL, (e) 6.69×10^4 m. **1.6** 2.36 lb. **1.7** 1.08×10^5 m^3. **1.8** 0.534 g/cm^3.

CHEMICAL

Mystery

The Disappearance of the Dinosaurs

Dinosaurs dominated life on Earth for millions of years and then disappeared very suddenly. To solve the mystery, paleontologists studied fossils and skeletons found in rocks in various layers of Earth's crust. Their findings enabled them to map out which species existed on Earth during specific geologic periods. They also revealed no dinosaur skeletons in rocks



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formed immediately after the Cretaceous period, which dates back some 65 million years. It is therefore assumed that the dinosaurs became extinct about 65 million years ago.

Among the many hypotheses put forward to account for their disappearance were disruptions of the food chain and a dramatic change in climate caused by violent volcanic eruptions. However, there was no convincing evidence for any one hypothesis until 1977. It was then that a group of paleontologists working in Italy obtained some very puzzling data at a site near Gubbio. The chemical analysis of a layer of clay deposited above sediments formed during the Cretaceous period (and therefore a layer that records events occurring *after* the Cretaceous period) showed a surprisingly high content of the element iridium (Ir). Iridium is very rare in Earth's crust but is comparatively abundant in asteroids.

This investigation led to the hypothesis that the extinction of dinosaurs occurred as follows. To account for the quantity of iridium found, scientists suggested that a large asteroid several miles in diameter hit Earth about the time the dinosaurs disappeared. The impact of the asteroid on Earth's surface must have been so tremendous that it literally vaporized a large quantity of surrounding rocks, soils, and other objects. The resulting dust and debris floated through the air and blocked the sunlight for months or perhaps years. Without ample sunlight most plants could not grow, and the fossil record confirms that many types of plants did indeed die out at this time. Consequently, of course, many plant-eating animals perished, and then, in turn, meat-eating animals began to starve. Dwindling food sources would obviously affect large animals needing great amounts of food more quickly and more severely than small animals. Therefore, the huge dinosaurs, the largest of which might have weighed as much as 30 tons, vanished due to lack of food.

Chemical Clues

1. How does the study of dinosaur extinction illustrate the scientific method?
2. Suggest two ways that would enable you to test the asteroid collision hypothesis.
3. In your opinion, is it justifiable to refer to the asteroid explanation as the theory of dinosaur extinction?
4. Available evidence suggests that about 20 percent of the asteroid's mass turned to dust and spread uniformly over Earth after settling out of the upper atmosphere. This dust amounted to about 0.02 g/cm^2 of Earth's surface. The asteroid very likely had a density of about 2 g/cm^3. Calculate the mass (in kilograms and tons) of the asteroid and its radius in meters, assuming that it was a sphere. (The area of Earth is 5.1×10^{14} m^2; 1 lb = 453.6 g,) (Source: *Consider a Spherical Cow—A Course in Environmental Problem Solving* by J. Harte, University Science Books, Mill Valley, CA 1988. Used with permission.)