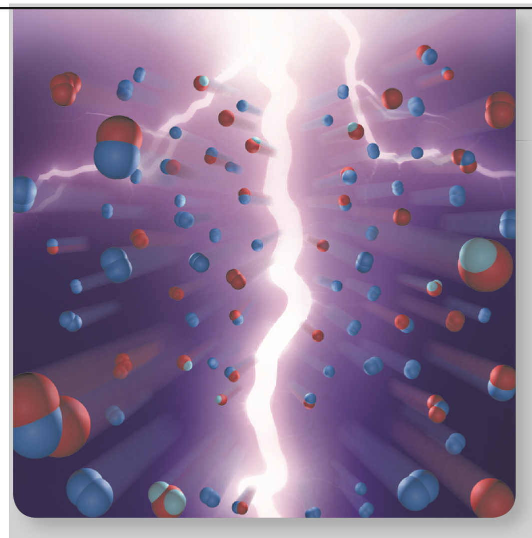
1 Keys to the Study of Chemistry

Key Principles
to focus on while studying this chapter

- *Matter* can undergo two kinds of change: *physical change* involves a change in *state*—gas, liquid, or solid—but not in ultimate makeup *(composition)*; *chemical change (reaction)* is more fundamental because it does involve a change in composition. The changes we observe result ultimately from changes too small to observe. ***(Section 1.1)***
- *Energy* occurs in different forms that are interconvertible, even as the total quantity of energy is conserved. When opposite charges are pulled apart, their *potential energy* increases; when they are released, potential energy is converted to the *kinetic energy* of the charges moving together. Matter consists of charged particles, so changes in energy accompany changes in matter. ***(Section 1.1)***
- The *scientific method* is a way of thinking that involves making *observations* and gathering *data* to develop *hypotheses* that are tested by *controlled experiments* until enough results are obtained to create a *model (theory)* that explains an aspect of nature. A sound theory can predict events but must be changed if new results conflict with it. ***(Section 1.2)***
- Any *measured quantity* is expressed by a number together with a *unit*. *Conversion factors* are ratios of equivalent quantities having different units; they are used in calculations to change the units of quantities. *Decimal prefixes* and *exponential notation* are used to express very large or very small quantities. ***(Section 1.3)***
- *The SI system* consists of seven *fundamental units*, each identifying a physical quantity such as length (meter), mass (kilogram), or temperature (kelvin). These are combined into many *derived units* used to identify quantities such as volume, density, and energy. *Extensive properties*, such as mass, depend on sample size; *intensive properties*, such as temperature, do not. ***(Section 1.4)***
- *Uncertainty* characterizes every measurement and is indicated by the number of *significant figures*. We *round* the final answer of a calculation to the same number of digits as in the least certain measurement. *Accuracy* refers to how close a measurement is to the true value; *precision* refers to how close measurements are to one another. ***(Section 1.5)***



A Molecular View Within a Storm Lightning supplies the energy for many atmospheric chemical changes to occur. In fact, all the events within and around you have causes and effects at the atomic level of reality.

Outline

CONCEPTS & SKILLS TO REVIEW
before studying this chapter

- exponential (scientific) notation (Appendix A)

Maybe you're taking this course because chemistry is fundamental to understanding other natural sciences. Maybe it's required for your major. Or maybe you just want to learn more about the impact of chemistry on society or even on your everyday life. For example, did you have cereal, fruit, and coffee for breakfast today? In chemical terms, you enjoyed nutrient-enriched, spoilage-retarded carbohydrate flakes mixed in a white emulsion of fats, proteins, and monosaccharides, with a piece of fertilizer-grown, pesticide-treated fruit, and a cup of hot aqueous extract of stimulating alkaloid. Earlier, you may have been awakened by the sound created as molecules aligned in the liquid-crystal display of your clock and electrons flowed to create a noise. You might have thrown off a thermal insulator of manufactured polymer and jumped in the shower to emulsify fatty substances on your skin and hair with purified water and formulated detergents. Perhaps you next adorned yourself in an array of pleasant-smelling pigmented gels, dyed polymeric fibers, synthetic footwear, and metal-alloy jewelry. After breakfast, you probably abraded your teeth with a colloidal dispersion of artificially flavored, dental-hardening agents, grabbed your laptop (an electronic device containing ultrathin, microetched semiconductor layers powered by a series of voltaic cells), collected some books (processed cellulose and plastic, electronically printed with light- and oxygen-resistant inks), hopped in your hydrocarbon-fueled, metal-vinyl-ceramic vehicle, electrically ignited a synchronized series of controlled gaseous explosions, and took off for class!

But the true impact of chemistry extends much farther than the products we use in daily life. The most profound questions about health, climate change, even the origin of life, ultimately have chemical answers.

No matter what your reason for studying chemistry, this course will help you develop two mental skills. The first, common to all science courses, is the ability to solve problems systematically. The second is specific to chemistry, for as you comprehend its ideas, you begin to view a hidden reality filled with incredibly minute particles moving at fantastic speeds, colliding billions of times a second, and interacting in ways that determine how all the matter inside and outside of you behaves. This chapter holds the keys to enter this world.

1.1 • SOME FUNDAMENTAL DEFINITIONS

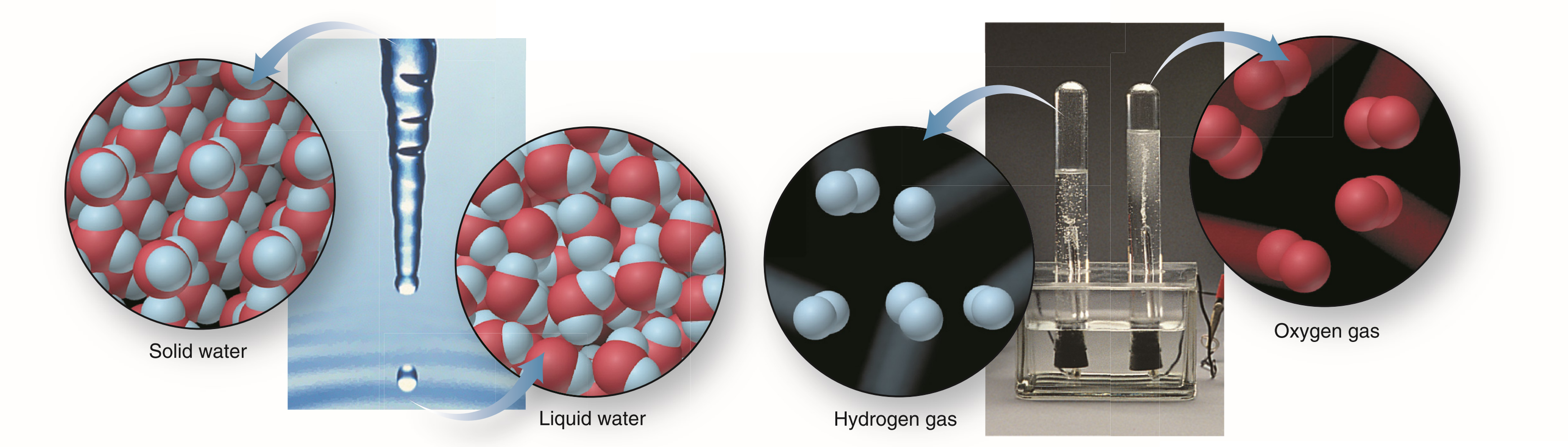
A good place to begin our exploration of chemistry is to define it and a few central concepts. **Chemistry** is *the study of matter and its properties, the changes that matter undergoes, and the energy associated with those changes.*

The Properties of Matter

Matter is the "stuff" of the universe: air, glass, planets, students—*anything that has mass and volume.* (In Section 1.4, we discuss the meanings of mass and volume in terms of how they are measured.) Chemists want to know the **composition** of matter, *the types and amounts of simpler substances that make it up.* A *substance* is a type of matter that has a defined, fixed composition.

We learn about matter by observing its **properties,** *the characteristics that give each substance its unique identity.* To identify a person, we might observe height, weight, hair and eye color, fingerprints, and even DNA pattern, until we arrive at a unique conclusion. To identify a substance, we observe two types of properties, *physical* and *chemical,* which are closely related to two types of change that matter undergoes:

- **Physical properties** are characteristics a substance shows *by itself, without changing into or interacting with another substance*. These properties include melting point, electrical conductivity, and density. A **physical change** occurs when a substance *alters its physical properties, **not** its composition.* For example, when ice melts,



A Physical change:
*Solid form of water becomes liquid form. Particles before and after remain the same, which means composition did **not** change.*

B Chemical change:
*Electric current decomposes water into different substances (hydrogen and oxygen). Particles before and after are different, which means composition **did** change.*

Figure 1.1 The distinction between physical and chemical change.

several physical properties change, such as hardness, density, and ability to flow. But the composition of the sample does *not* change: it is still water. The photograph in Figure 1.1A shows what this change looks like in everyday life. The "blow-up" circles depict a magnified view of the particles making up the sample. In the icicle, the particles lie in a repeating pattern, whereas they are jumbled in the droplet, but *the particles are the same* in both forms of water.

Physical change (same substance before and after):

$$\text{Water (solid form)} \longrightarrow \text{water (liquid form)}$$

- **Chemical properties** are characteristics a substance shows *as it changes into or interacts with another substance (or substances)*. Chemical properties include flammability, corrosiveness, and reactivity with acids. A **chemical change,** also called a **chemical reaction,** occurs when *a substance (or substances) is converted into a different substance (or substances)*. Figure 1.1B shows the chemical change (reaction) that occurs when you pass an electric current through water: the water decomposes (breaks down) into two other substances, hydrogen and oxygen, that bubble into the tubes. The composition *has* changed: the final sample is no longer water.

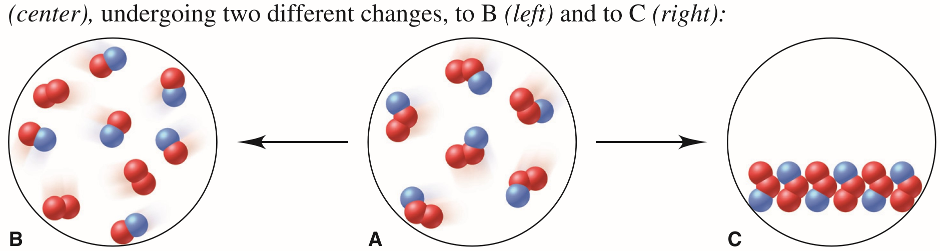
  Chemical change (different substances before and after):

$$\text{Water} \xrightarrow{\text{electric current}} \text{hydrogen} + \text{oxygen}$$

Let's work through a sample problem that uses atomic-scale scenes to distinguish between physical and chemical change.

Sample Problem 1.1 Visualizing Change on the Atomic Scale

Problem The scenes below represent an atomic-scale view of a sample of matter, A *(center)*, undergoing two different changes, to B *(left)* and to C *(right)*:

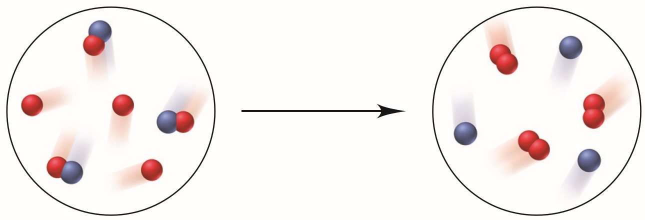


Decide whether each depiction shows a physical or a chemical change.

Plan Given depictions of two changes, we have to determine whether each represents a physical or a chemical change. The number and colors of the little spheres that make up each particle tell its "composition." Samples with particles of the *same* composition but in a different arrangement depict a *physical* change, whereas samples with particles of a *different* composition depict a *chemical* change.

Solution In A, each particle consists of one blue and two red spheres. The particles in A change into two types in B, one made of red and blue spheres and the other made of two red spheres; therefore, they have undergone a chemical change to form different particles. The particles in C are the same as those in A, but they are closer together and arranged differently; therefore, they have undergone a physical change.

FOLLOW-UP PROBLEM 1.1 Is the following change chemical or physical? (Compare your answer with the one in Brief Solutions to Follow-up Problems at the end of the chapter.)



The States of Matter

Matter occurs commonly in *three physical forms* called **states:** solid, liquid, and gas. We'll define the states and see how temperature can change them.

Defining the States On the macroscopic scale, each state of matter is defined by the way the sample fills a container (Figure 1.2, *flasks at top*):

- A **solid** has a fixed shape that does not conform to the container shape. Solids are *not* defined by rigidity or hardness: solid iron is rigid and hard, but solid lead is flexible, and solid wax is soft.
- A **liquid** has a varying shape that conforms to the container shape, but only to the extent of the liquid's volume; that is, a liquid has *an upper surface.*
- A **gas** also has a varying shape that conforms to the container shape, but it fills the entire container and, thus, does *not* have a surface.

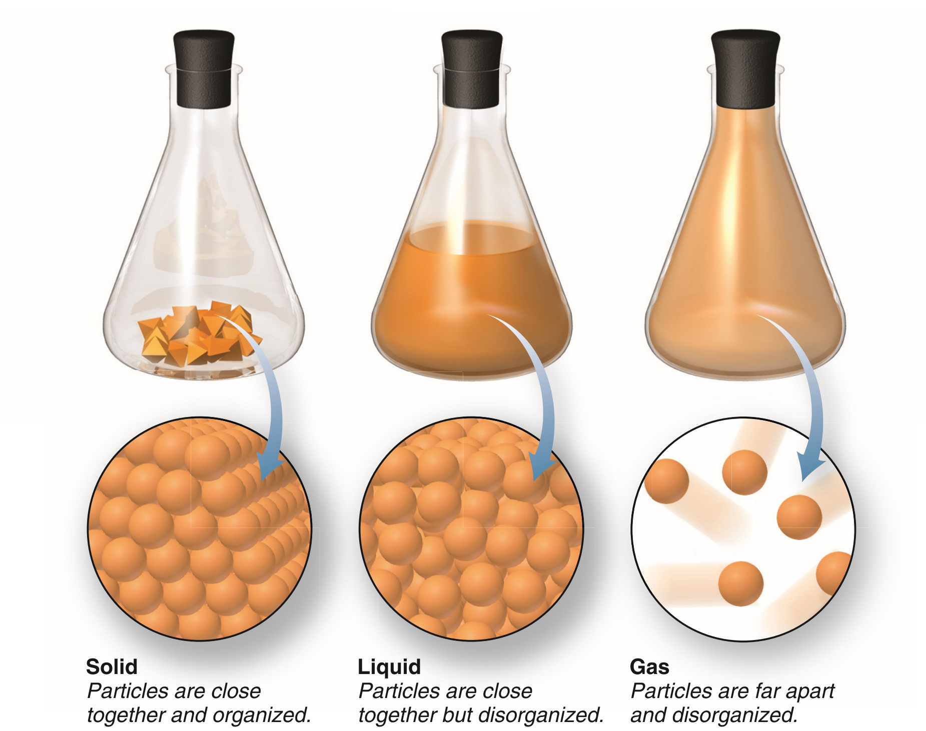


Figure 1.2 The physical states of matter.

On the atomic scale, each state is defined by the relative positions of its particles (Figure 1.2, *circles at bottom*):

- In a *solid,* the particles lie next to each other in a regular, three-dimensional *array.*
- In a *liquid,* the particles also lie close together but move randomly around each other.
- In a *gas,* the particles have large distances between them and move randomly throughout the container.

Temperature and Changes of State Depending on the temperature and pressure of the surroundings, many substances can exist in each of the three physical states and undergo changes in state as well. For example, as the temperature increases, solid water melts to liquid water, which boils to gaseous water (also called *water vapor*). Similarly, as the temperature drops, water vapor condenses to liquid water, and with further cooling, the liquid freezes to ice. The majority of other substances—such as benzene, nitrogen, and iron—can undergo similar changes of state.

The main point is that *a physical change caused by heating can generally be reversed by cooling*. This is *not* generally true for a chemical change. For example, heating iron in moist air causes a chemical reaction that yields the brown, crumbly substance known as rust. Cooling does not reverse this change; rather, another chemical change (or series of them) is required.

The following sample problem provides practice in distinguishing some familiar examples of physical and chemical change.

Sample Problem 1.2 Distinguishing Between Physical and Chemical Change

Problem Decide whether each of the following processes is primarily a physical or a chemical change, and explain briefly:
(a) Frost forms as the temperature drops on a humid winter night.
(b) A cornstalk grows from a seed that is watered and fertilized.
(c) A match ignites to form ash and a mixture of gases.
(d) Perspiration evaporates when you relax after jogging.
(e) A silver fork tarnishes slowly in air.

Plan To decide whether a change is chemical or physical, we ask, "Does the substance change composition or just change form?"

Solution **(a)** Frost forming is a physical change: the drop in temperature changes water vapor (gaseous water) in humid air to ice crystals (solid water).
(b) A seed growing involves chemical change: the seed uses water, substances from air, fertilizer, and soil, and energy from sunlight to make complex changes in composition.
(c) The match burning is a chemical change: the combustible substances in the match head are converted into other substances.
(d) Perspiration evaporating is a physical change: the water in sweat changes its form, from liquid to gas, but not its composition.
(e) Tarnishing is a chemical change: silver changes to silver sulfide by reacting with sulfur-containing substances in the air.

FOLLOW-UP PROBLEM 1.2 Is each of the following processes primarily a physical or a chemical change? Explain. (See Brief Solutions at the end of the chapter.)
(a) Purple iodine vapor appears when solid iodine is warmed.
(b) Gasoline fumes are ignited by a spark in an automobile engine's cylinder.
(c) A scab forms over an open cut.

The Central Theme in Chemistry

Understanding the properties of a substance and the changes it undergoes leads to the central theme in chemistry: *macroscopic-scale* properties and behavior, those we can see, are the results of *atomic-scale* properties and behavior that we cannot see. The distinction between chemical and physical change is defined by composition, which we study macroscopically. But composition ultimately depends on the makeup of substances at the atomic scale. Similarly, macroscopic properties of substances in any of the three states arise from atomic-scale behavior of their particles. Picturing a chemical event on the molecular scale, even one as common as the flame of a candle, helps clarify what is taking place. What is happening when water boils or copper melts? What events occur in the invisible world of minute particles that cause a seed to grow, a neon light to glow, or a nail to rust? Throughout the text, we return to this central idea: *we study **observable** changes in matter to understand their **unobservable** causes.*

The Importance of Energy in the Study of Matter

Physical and chemical changes are accompanied by energy changes. **Energy** is often defined as *the ability to do work.* Essentially, all work involves moving something. Work is done when your arm lifts a book, when a car's engine moves the wheels, or when a falling rock moves the ground as it lands. The object doing the work (arm, engine, rock) transfers some of its energy to the object on which the work is done (book, wheels, ground).

The total energy an object possesses is the sum of its potential energy and its kinetic energy.

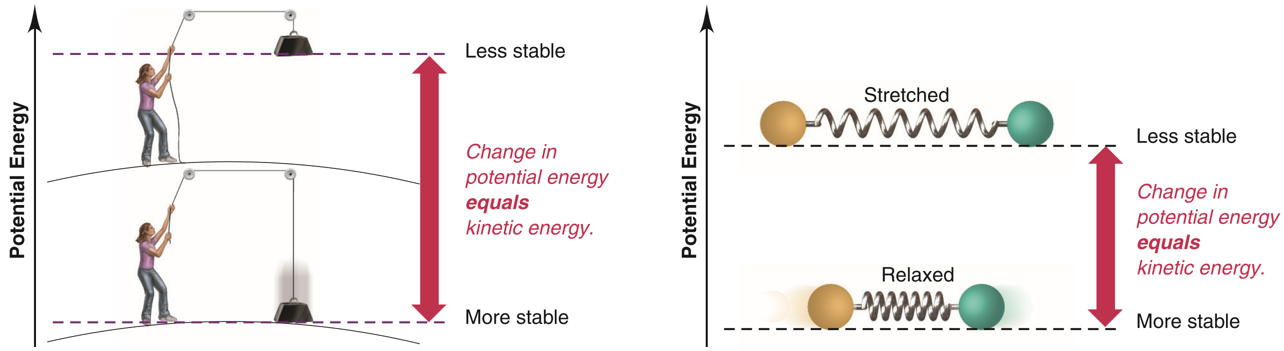
- **Potential energy** is *the energy due to the **position** of the object relative to other objects.*
- **Kinetic energy** is *the energy due to the **motion** of the object.*

Let's examine four systems that illustrate the relationship between these two forms of energy: a weight raised above the ground, two balls attached by a spring, two electrically charged particles, and a burning fuel and its waste products. Two concepts central to all these cases are

1. *When energy is converted from one form to the other, it is conserved, not destroyed.*
2. *Situations of lower energy are more stable, and therefore favored, over situations of higher energy (less stable).*

The four systems are depicted in Figure 1.3 on the next page:

- *A weight raised above the ground* (Figure 1.3A). The energy you exert to lift a weight against gravity increases the weight's potential energy (energy due to its position). When you drop the weight, that additional potential energy is converted to kinetic energy (energy due to motion). The situation with the weight elevated and higher in potential energy is *less stable,* so the weight will fall when released to result in a situation that is lower in potential energy and *more stable.*
- *Two balls attached by a spring* (Figure 1.3B). When you pull the balls apart, the energy you exert to stretch the relaxed spring increases the system's potential energy. This change in potential energy is converted to kinetic energy when you release the balls. The system of balls and spring is less stable (has more potential energy) when the spring is stretched than when it is relaxed.
- *Two electrically charged particles* (Figure 1.3C). Due to interactions known as *electrostatic forces, opposite charges attract each other, and like charges repel each other.* When energy is exerted to move a positive particle away from a negative one, the potential energy of the system increases, and that increase is converted to kinetic energy when the particles are pulled together by the electrostatic attraction. Similarly, when energy is used to move two positive (or two negative) particles together, their potential energy increases and changes to kinetic energy when they are pushed apart by the electrostatic repulsion. Charged particles move naturally to a more stable situation (lower energy).
- *A burning fuel and its waste products* (Figure 1.3D). Matter is composed of positively and negatively charged particles. *The chemical potential energy of a substance results from the relative positions of and the attractions and repulsions among its*



A **A gravitational system.** *Potential energy is gained when a weight is lifted. It is converted to kinetic energy as the weight falls.*

B **A system of two balls attached by a spring.** *Potential energy is gained when the spring is stretched. It is converted to the kinetic energy of the moving balls as the spring relaxes.*

C **A system of oppositely charged particles.** *Potential energy is gained when the charges are separated. It is converted to kinetic energy as the attraction pulls the charges together.*

D **A system of fuel and exhaust.** *A fuel is higher in chemical potential energy than the exhaust. As the fuel burns, some of its potential energy is converted to the kinetic energy of the moving car.*

Figure 1.3 Potential energy is converted to kinetic energy. The dashed horizontal lines indicate the potential energy of each system before and after the change.

particles. Some substances are higher in potential energy than others. For example, gasoline and oxygen have more chemical potential energy than the exhaust gases they form. This difference is converted into kinetic energy, which moves the car, heats the interior, makes the lights shine, and so on. Similarly, the difference in potential energy between the food and air we take in and the wastes we excrete enables us to move, grow, keep warm, study chemistry, and so on.

Summary of Section 1.1

- Chemists study the composition and properties of matter and how they change.
- Each substance has a unique set of *physical* properties (attributes of the substance itself) and *chemical* properties (attributes of the substance as it interacts with or changes to other substances). Changes in matter can be *physical* (different form of the same substance) or *chemical* (different substance).
- Matter exists in three physical states—solid, liquid, and gas. The behavior of each state is due to the arrangement of the particles.
- A physical change caused by heating may be reversed by cooling. But a chemical change caused by heating can be reversed only by other chemical changes.
- Macroscopic changes result from submicroscopic changes.
- Changes in matter are accompanied by changes in energy.
- An object's potential energy is due to its position; an object's kinetic energy is due to its motion. Energy used to lift a weight, stretch a spring, or separate opposite charges increases the system's potential energy, which is converted to kinetic energy as the system returns to its original condition. Energy changes form but is conserved.
- Chemical potential energy arises from the positions and interactions of a substance's particles. When a higher energy (less stable) substance is converted into a more stable (lower energy) substance, some potential energy is converted into kinetic energy.

1.2 • THE SCIENTIFIC APPROACH: DEVELOPING A MODEL

Unlike our prehistoric ancestors, who survived through *trial and error*—gradually learning which types of stone were hard enough to shape others, which plants were edible and which poisonous—we employ the *quantitative theories* of chemistry to understand materials, make better use of them, and create new ones: specialized drugs to target diseases, advanced composites for vehicles, synthetic polymers for clothing and sports gear, liquid crystals for electronic displays, and countless others.

To understand nature, scientists use an approach called the **scientific method.** It is not a stepwise checklist, but rather a process involving creative propositions and tests aimed at objective, verifiable discoveries. There is no single procedure, and luck often plays a key role in discovery. In general terms, the scientific approach includes the following parts (Figure 1.4):

- ***Observations.*** These are the facts our ideas must explain. The most useful observations are quantitative because they can be analyzed to reveal trends. Pieces of quantitative information are **data.** When the same observation is made by many investigators in many situations with no clear exceptions, it is summarized, often in mathematical terms, as a **natural law.** The observation that mass remains constant during chemical change—made in the 18th century by the French chemist Antoine Lavoisier (1743–1794) and numerous experimenters since—is known as the law of mass conservation (Chapter 2).
- ***Hypothesis.*** Whether derived from observation or from a "spark of intuition," a hypothesis is a proposal made to explain an observation. A sound hypothesis need not be correct, but it must be *testable by experiment.* Indeed, a hypothesis is often the reason for performing an experiment: if the results do not support it, the hypothesis must be revised or discarded. Hypotheses can be altered, but experimental results cannot.
- ***Experiment.*** A set of procedural steps that tests a hypothesis, an experiment often leads to a revised hypothesis and new experiments to test it. An experiment typically contains at least two **variables,** quantities that can have more than one value. A well-designed experiment is **controlled** in that it measures the effect of one variable on another while keeping all other variables constant. Experimental results must be *reproducible* by others. Both skill and creativity play a part in experimental design.
- ***Model.*** Formulating conceptual models, or **theories,** based on *experiments* that test *hypotheses* about *observations* distinguishes scientific thinking from speculation. As hypotheses are revised according to experimental results, a model emerges to explain how the phenomenon occurs. A model is a *simplified,* not an exact, representation of some aspect of nature that we use to *predict* related phenomena. Ongoing experimentation refines the model to account for new facts.

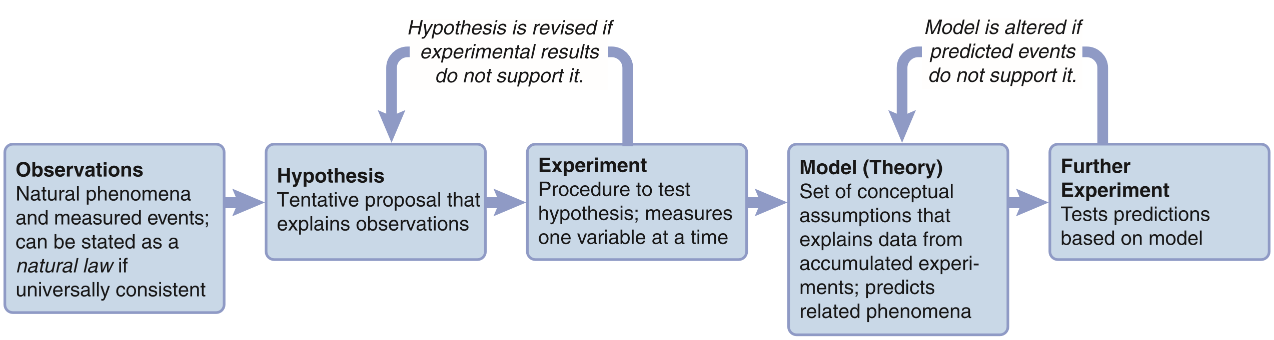


Figure 1.4 The scientific approach to understanding nature. Hypotheses and models are mental pictures that are revised to match observations and experimental results, *not* the other way around.

The following short paragraph is the first of an occasional feature that will help you learn a concept through an analogy, a unifying idea, or a memorization aid.

THINK OF IT THIS WAY
Everyday Scientific Thinking

Consider this familiar scenario. While listening to an FM station on your car's audio system, you notice the sound is garbled (observation) and assume it is caused by poor reception (hypothesis). To isolate this variable, you plug in your MP3 player and listen to a song (experiment): the sound is still garbled. If the problem is not poor reception, perhaps the speakers are at fault (new hypothesis). To isolate this variable, you listen with headphones (experiment): the sound is clear. You conclude that the speakers need to be repaired (model). The repair shop says the speakers are fine (new observation), but the car's amplifier may be at fault (new hypothesis). Repairing the amplifier corrects the garbled sound (new experiment), so the amplifier was the problem (revised model). Approaching a problem scientifically is a common practice, even if you're not aware of it.

Summary of Section 1.2

- The scientific method is a process designed to explain and predict phenomena.
- Observations lead to hypotheses about how or why a phenomenon occurs. When repeated with no exceptions, observations may be expressed as a natural law.
- Hypotheses are tested by controlled experiments and revised when necessary.
- If reproducible data support a hypothesis, a model (theory) can be developed to explain the observed phenomenon. A good model predicts related phenomena but must be refined whenever conflicting data appear.

1.3 • CHEMICAL PROBLEM SOLVING

In many ways, learning chemistry is learning how to solve chemistry problems. This section describes the problem-solving approach used throughout this book. Most problems include calculations, so let's first discuss how to handle measured quantities.

Units and Conversion Factors in Calculations

All measured quantities consist of a number *and* a unit: a person's height is "5 feet, 10 inches," not "5, 10." Ratios of quantities have ratios of units, such as miles/hour. (We discuss some important units in Section 1.4.) To minimize errors, make it a habit to *include units in all calculations*.

The arithmetic operations used with quantities are the same as those used with pure numbers; that is, units can be multiplied, divided, and canceled:

- A carpet measuring 3 feet by 4 feet (ft) has an area of

$$\text{Area} = 3\text{ ft} \times 4\text{ ft} = (3 \times 4)(\text{ft} \times \text{ft}) = 12\text{ ft}^2$$

- A car traveling 350 miles (mi) in 7 hours (h) has a speed of

$$\text{Speed} = \frac{350\text{ mi}}{7\text{ h}} = \frac{50\text{ mi}}{1\text{ h}}\ (\text{often written } 50\text{ mi}\cdot\text{h}^{-1})$$

- In 3 hours, the car travels a distance of

$$\text{Distance} = 3\ \cancel{\text{h}} \times \frac{50\text{ mi}}{1\ \cancel{\text{h}}} = 150\text{ mi}$$

Constructing a Conversion Factor **Conversion factors** are *ratios used to express a quantity in different units.* Suppose we want to know the distance of that 150-mile car trip in feet. To convert miles to feet, we use *equivalent quantities,*

$$1\text{ mi} = 5280\text{ ft}$$

from which we can construct two conversion factors. Dividing both sides by 5280 ft gives one conversion factor (shown in blue):

$$\frac{1\ \text{mi}}{5280\ \text{ft}} = \frac{5280\ \cancel{\text{ft}}}{5280\ \cancel{\text{ft}}} = 1$$

And, dividing both sides by 1 mi gives the other conversion factor (the inverse):

$$\frac{1\ \cancel{\text{mi}}}{1\ \cancel{\text{mi}}} = \frac{5280\ \text{ft}}{1\ \text{mi}} = 1$$

Since the numerator and denominator of a conversion factor are equal, multiplying a quantity by a conversion factor is the same as multiplying by 1. Thus, *even though the number and unit change, the size of the quantity remains the same.*

To convert the distance from miles to feet, we choose the conversion factor with miles in the denominator, because it cancels miles and gives the answer in feet:

$$\text{Distance (ft)} = 150\ \cancel{\text{mi}} \times \frac{5280\ \text{ft}}{1\ \cancel{\text{mi}}} = 792{,}000\ \text{ft}$$

$$\text{mi} \quad \Rightarrow \quad \text{ft}$$

Choosing the Correct Conversion Factor It is easier to convert if you first decide whether the answer expressed in the new units should have a larger or smaller number. In the previous case, we know that a foot is *smaller* than a mile, so the distance in feet should have a *larger* number (792,000) than the distance in miles (150). The conversion factor has the larger number (5280) in the numerator, so it gave a larger number in the answer.

Most importantly, the *conversion factor you choose must cancel all units except those you want in the answer.* Therefore, set the unit you are converting *from* (beginning unit) in the *opposite position in the conversion factor* (numerator or denominator) so that it cancels and you are left with the unit you are converting *to* (final unit):

$$\cancel{\text{beginning unit}} \times \frac{\text{final unit}}{\cancel{\text{beginning unit}}} = \text{final unit} \qquad \text{as in} \qquad \cancel{\text{mi}} \times \frac{\text{ft}}{\cancel{\text{mi}}} = \text{ft}$$

Or, in cases that involve units raised to a power:

$$(\cancel{\text{beginning unit}} \times \cancel{\text{beginning unit}}) \times \frac{\text{final unit}^2}{\cancel{\text{beginning unit}^2}} = \text{final unit}^2$$

$$\text{as in} \qquad (\cancel{\text{ft}} \times \cancel{\text{ft}}) \times \frac{\text{mi}^2}{\cancel{\text{ft}^2}} = \text{mi}^2$$

Or, in cases that involve a ratio of units:

$$\frac{\cancel{\text{beginning unit}}}{\text{final unit}_1} \times \frac{\text{final unit}_2}{\cancel{\text{beginning unit}}} = \frac{\text{final unit}_2}{\text{final unit}_1} \qquad \text{as in} \qquad \frac{\cancel{\text{mi}}}{\text{h}} \times \frac{\text{ft}}{\cancel{\text{mi}}} = \frac{\text{ft}}{\text{h}}$$

Converting Between Unit Systems We use the same procedure to convert between systems of units, for example, between the English (or American) unit system and the International System (a revised metric system; Section 1.4). Suppose we know that the height of Angel Falls in Venezuela (the world's highest) is 3212 ft, and we find its height in miles as

$$\text{Height (mi)} = 3212\ \cancel{\text{ft}} \times \frac{1\ \text{mi}}{5280\ \cancel{\text{ft}}} = 0.6083\ \text{mi}$$

$$\text{ft} \quad \Rightarrow \quad \text{mi}$$

Now, we want its height in kilometers (km). The equivalent quantities are

$$1.609\ \text{km} = 1\ \text{mi}$$

Because we are converting *from* miles *to* kilometers, we use the conversion factor with miles in the denominator in order to cancel miles:

$$\text{Height (km)} = 0.6083\ \cancel{\text{mi}} \times \frac{1.609\ \text{km}}{1\ \cancel{\text{mi}}} = 0.9788\ \text{km}$$

$$\text{mi} \quad \Rightarrow \quad \text{km}$$

Notice that kilometers are *smaller* than miles, so this conversion factor gave us an answer with a *larger* number (0.9788 is larger than 0.6083).

If we want the height of Angel Falls in meters (m), we use the equivalent quantities 1 km = 1000 m to construct the conversion factor:

$$\text{Height (m)} = 0.9788\ \cancel{\text{km}} \times \frac{1000\ \text{m}}{1\ \cancel{\text{km}}} = 978.8\ \text{m}$$

$$\text{km} \Rightarrow \text{m}$$

In longer calculations, we often string together several conversion steps:

$$\text{Height (m)} = 3212\ \cancel{\text{ft}} \times \frac{1\ \cancel{\text{mi}}}{5280\ \cancel{\text{ft}}} \times \frac{1.609\ \cancel{\text{km}}}{1\ \cancel{\text{mi}}} \times \frac{1000\ \text{m}}{1\ \cancel{\text{km}}} = 978.8\ \text{m}$$

$$\text{ft} \Rightarrow \text{mi} \Rightarrow \text{km} \Rightarrow \text{m}$$

A Systematic Approach to Solving Chemistry Problems

The approach used in this book to solve problems emphasizes reasoning, not memorizing, and is based on a simple idea: plan how to solve the problem *before* you try to solve it, then check your answer, and practice with a similar follow-up problem. In general, the sample problems consist of several parts:

1. **Problem.** This part states all the information you need to solve the problem, usually framed in some interesting context.
2. **Plan.** This part helps you *think* about the solution *before* juggling numbers and pressing calculator buttons. There is often more than one way to solve a problem, and the given plan is one possibility. The plan will
   - Clarify the known and unknown: what information do you have, and what are you trying to find?
   - Suggest the steps from known to unknown: what ideas, conversions, or equations are needed?
   - Present a road map (especially in early chapters), a flow diagram of the plan. The road map has a box for each intermediate result and an arrow showing the step (conversion factor or operation) used to get to the next box.
3. **Solution.** This part shows the calculation steps in the same order as in the plan (and the road map).
4. **Check.** This part helps you check that your final answer makes sense: Are the units correct? Did the change occur in the expected direction? Is it reasonable chemically? To avoid a large math error, we also often do a rough calculation and see if we get an answer "in the same ballpark" as the actual result. Here's a typical "ballpark" calculation from everyday life. You are at a clothing store and buy three shirts at \$14.97 each. With a 5% sales tax, the bill comes to \$47.16. In your mind, you know that \$14.97 is about \$15, and 3 times \$15 is \$45; with the sales tax, the cost should be a bit more. So, your quick mental calculation *is* in the same ballpark as the actual cost.
5. **Comment.** This part appears occasionally to provide an application, an alternative approach, a common mistake to avoid, or an overview.
6. **Follow-up Problem.** This part presents a similar problem that requires you to apply concepts and/or methods used in solving the sample problem.

Of course, you can't learn to solve chemistry problems, any more than you can learn to swim, by reading about it, so here are a few suggestions:

- Follow along in the sample problem with pencil, paper, and calculator.
- Try the follow-up problem as soon as you finish the sample problem. A feature called Brief Solutions to Follow-up Problems appears at the end of each chapter, allowing you to compare your solution steps and answer.
- Read the sample problem and text again if you have trouble.

- Go to the Connect website for this text at www.mcgrawhillconnect.com and do the homework assignment. Hints and feedback on common incorrect answers, as well as step-by-step solutions, will help you learn to be an effective problem solver.
- The end-of-chapter problems review and extend the concepts and skills in the chapter, so work as many as you can. (Answers are given in the back of the book for problems with a colored number.)

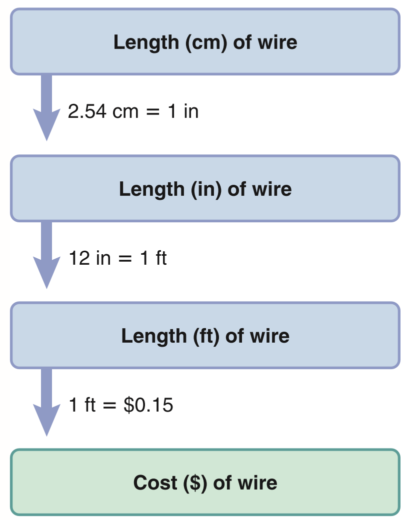
Let's apply this systematic approach in a unit-conversion problem.

Sample Problem 1.3 Converting Units of Length

Problem To hang some paintings in your dorm, you need 325 centimeters (cm) of picture wire that sells for \$0.15/ft. How much does the wire cost?

Plan We know the length of wire in centimeters (325 cm) and the price in dollars per foot (\$0.15/ft). We can find the unknown cost of the wire by converting the length from centimeters to inches (in) and from inches to feet. The price gives us the equivalent quantities (1 ft = \$0.15) to convert feet of wire to cost in dollars. The road map starts with the known and moves through the calculation steps to the unknown.

Road Map



Solution Converting the known length from centimeters to inches: The equivalent quantities alongside the road map arrow are needed to construct the conversion factor. We choose 1 in/2.54 cm, rather than the inverse, because it gives an answer in inches:

$$\text{Length (in)} = \text{length (cm)} \times \text{conversion factor} = 325\ \cancel{\text{cm}} \times \frac{1\ \text{in}}{2.54\ \cancel{\text{cm}}} = 128\ \text{in}$$

Converting the length from inches to feet:

$$\text{Length (ft)} = \text{length (in)} \times \text{conversion factor} = 128\ \cancel{\text{in}} \times \frac{1\ \text{ft}}{12\ \cancel{\text{in}}} = 10.7\ \text{ft}$$

Converting the length in feet to cost in dollars:

$$\text{Cost (\$)} = \text{length (ft)} \times \text{conversion factor} = 10.7\ \cancel{\text{ft}} \times \frac{\$0.15}{1\ \cancel{\text{ft}}} = \boxed{\$1.60}$$

Check The units are correct for each step. The conversion factors make sense in terms of the relative unit sizes: the number of inches is *smaller* than the number of centimeters (an inch is *larger* than a centimeter), and the number of feet is *smaller* than the number of inches. The total cost seems reasonable: a little more than 10 ft of wire at \$0.15/ft should cost a little more than \$1.50.

Comment **1.** We could also have strung the three steps together:

$$\text{Cost (\$)} = 325\ \cancel{\text{cm}} \times \frac{1\ \cancel{\text{in}}}{2.54\ \cancel{\text{cm}}} \times \frac{1\ \cancel{\text{ft}}}{12\ \cancel{\text{in}}} \times \frac{\$0.15}{1\ \cancel{\text{ft}}} = \boxed{\$1.60}$$

2. There are usually alternative sequences in unit-conversion problems. Here, for example, we would get the same answer if we first converted the cost of wire from \$/ft to \$/cm and kept the wire length in cm. Try it yourself.

FOLLOW-UP PROBLEM 1.3 A furniture factory needs 31.5 ft^2 of fabric to upholster one chair. Its Dutch supplier sends the fabric in bolts of exactly 200 m^2. How many chairs can be upholstered with 3 bolts of fabric (1 m = 3.281 ft)? Draw a road map to show how you plan the solution. (See Brief Solutions.)

Summary of Section 1.3

- A measured quantity consists of a number and a unit.
- A conversion factor is a ratio of equivalent quantities (and, thus, equal to 1) that is used to express a quantity in different units.

- The problem-solving approach used in this book has four parts: (1) plan the steps to the solution, which often includes a flow diagram (road map) of the steps, (2) perform the calculations according to the plan, (3) check to see if the answer makes sense, and (4) practice with a similar problem and compare your solution with the one at the end of the chapter.

1.4 • MEASUREMENT IN SCIENTIFIC STUDY

Almost everything we own is made and sold in measured amounts. The measurement systems we use have a rich history characterized by the search for *exact, invariable standards.* Measuring for purposes of trade, building, and surveying used to be based on standards that could vary: a yard was the distance from the king's nose to the tip of his outstretched arm, and an acre was the area tilled in one day by a man with a pair of oxen. Our current, far more exact system of measurement began in 1790 when a committee in France developed the original *metric system.* In 1960, another committee in France revised it to create the universally accepted **SI units** (from the French **S**ystème **I**nternational d'Unités).

General Features of SI Units

The SI system is based on seven **fundamental units,** or **base units,** each identified with a physical quantity (Table 1.1). All other units are **derived units,** combinations of the seven base units. For example, the derived unit for speed, meters per second (m/s), is the base unit for length (m) divided by the base unit for time (s). (Derived units that are a *ratio* of base units can be used as conversion factors.) For quantities much smaller or larger than the base unit, we use decimal prefixes and exponential (scientific) notation (Table 1.2). (If you need a review of exponential notation, see Appendix A.) Because the prefixes are based on powers of 10, SI units are easier to use in calculations than English units.

Table 1.1 SI Base Units

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

Some Important SI Units in Chemistry

Here, we discuss units for length, volume, mass, density, temperature, and time; other units are presented in later chapters. Table 1.3 shows some SI quantities for length, volume, and mass, along with their English-system equivalents.

Length The SI base unit of length is the **meter (m),** which is about 2.5 times the width of this book when open. The definition is exact and invariant: 1 meter is the distance light travels in a vacuum in 1/299,792,458 of a second. A meter is a little longer than a yard (1 m = 1.094 yd); a centimeter (10^{-2} m) is about two-fifths of an inch (1 cm = 0.3937 in; 1 in = 2.54 cm). Biological cells are often measured in micrometers (1 μm = 10^{-6} m). On the atomic scale, nanometers (10^{-9} m) and picometers (10^{-12} m) are used. Many proteins have diameters of about 2 nm; atomic diameters are about 200 pm (0.2 nm). An older unit still in use is the angstrom (1 Å = 10^{-10} m = 0.1 nm = 100 pm).

Table 1.2 Common Decimal Prefixes Used with SI Units

Prefix*	Prefix Symbol	Word	Conventional Notation	Exponential Notation
tera	T	trillion	1,000,000,000,000	1×10^{12}
giga	G	billion	1,000,000,000	1×10^{9}
mega	M	million	1,000,000	1×10^{6}
kilo	k	thousand	1,000	1×10^{3}
hecto	h	hundred	100	1×10^{2}
deka	da	ten	10	1×10^{1}
—	—	one	1	1×10^{0}
deci	d	tenth	0.1	1×10^{-1}
centi	c	hundredth	0.01	1×10^{-2}
milli	m	thousandth	0.001	1×10^{-3}
micro	μ	millionth	0.000001	1×10^{-6}
nano	n	billionth	0.000000001	1×10^{-9}
pico	p	trillionth	0.000000000001	1×10^{-12}
femto	f	quadrillionth	0.000000000000001	1×10^{-15}

*The prefixes most frequently used by chemists appear in bold type.

Volume Any sample of matter has a certain **volume (*V*),** the amount of space it occupies. The SI unit of volume is the **cubic meter (m^3).** In chemistry, we often use the non-SI units **liter (L)** and **milliliter (mL)** (note the uppercase L). Medical practitioners measure body fluids in cubic decimeters (dm^3), which are equivalent to liters:

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

And 1 mL, or $\frac{1}{1000}$ of a liter, is equivalent to 1 cubic centimeter (cm^3):

$$1\ \text{mL} = 1\ \text{cm}^3 = 10^{-3}\ \text{dm}^3 = 10^{-3}\ \text{L} = 10^{-6}\ \text{m}^3$$

A liter is slightly larger than a quart (qt) (1 L = 1.057 qt; 1 qt = 946.4 mL); 1 fluid ounce ($\frac{1}{32}$ of a quart) equals 29.57 mL (29.57 cm^3).

Figure 1.5 shows some laboratory glassware for working with volumes. Volumetric flasks and pipets have a fixed volume indicated by a mark on the neck.

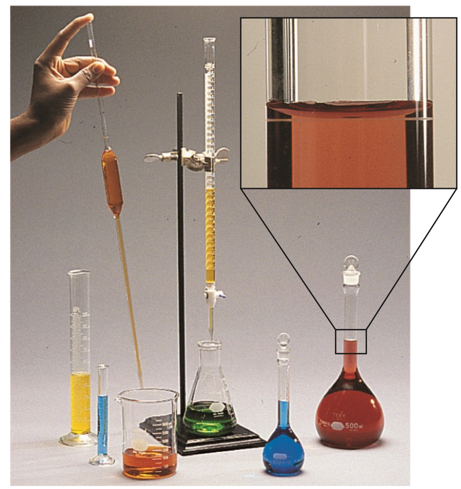


Figure 1.5 Common laboratory volumetric glassware. From left to right are two graduated cylinders, a pipet being emptied into a beaker, a buret delivering liquid to an Erlenmeyer flask, and two volumetric flasks. **Inset,** In contact with the glass neck, the liquid forms a concave meniscus (curved surface).

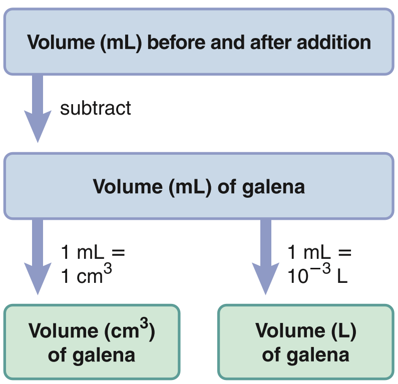
Table 1.3 Common SI-English Equivalent Quantities

Quantity	SI	SI Equivalents	English Equivalents	English to SI Equivalent
Length	1 kilometer (km)	1000 (10^3) meters	0.6214 mile (mi)	1 mile = 1.609 km
	1 meter (m)	100 (10^2) centimeters	1.094 yards (yd)	1 yard = 0.9144 m
		1000 millimeters (mm)	39.37 inches (in)	1 foot (ft) = 0.3048 m
	1 centimeter (cm)	0.10 (10^{-2}) meter	0.3937 inch	1 inch = 2.54 cm (exactly)
Volume	1 cubic meter (m^3)	1,000,000 (10^6) cubic centimeters	35.31 cubic feet (ft^3)	1 cubic foot = 0.02832 m^3
	1 cubic decimeter (dm^3)	1000 cubic centimeters	0.2642 gallon (gal)	1 gallon = 3.785 dm^3
			1.057 quarts (qt)	1 quart = 0.9464 dm^3
				1 quart = 946.4 cm^3
	1 cubic centimeter (cm^3)	0.001 dm^3	0.03381 fluid ounce	1 fluid ounce = 29.57 cm^3
Mass	1 kilogram (kg)	1000 grams	2.205 pounds (lb)	1 pound = 0.4536 kg
	1 gram (g)	1000 milligrams (mg)	0.03527 ounce (oz)	1 ounce = 28.35 g

Sample Problem 1.4 Converting Units of Volume

Problem The volume of an irregularly shaped solid can be determined from the volume of water it displaces. A graduated cylinder contains 19.9 mL of water. When a small piece of galena, an ore of lead, is added, it sinks and the volume increases to 24.5 mL. What is the volume of the piece of galena in cm^3 and in L?

Road Map



Plan We have to find the volume of the galena from the change in volume of the cylinder contents. The volume of galena in mL is the difference before (19.9 mL) and after (24.5 mL) adding it. Since mL and cm^3 represent identical volumes, the volume in mL equals the volume in cm^3. We then use equivalent quantities (1 mL $= 10^{-3}$ L) to convert mL to L. The road map shows these steps.

Solution Finding the volume of galena:

$$\text{Volume (mL)} = \text{volume after} - \text{volume before} = 24.5\ \text{mL} - 19.9\ \text{mL} = 4.6\ \text{mL}$$

Converting the volume from mL to cm^3:

$$\text{Volume (cm}^3\text{)} = 4.6\ \cancel{\text{mL}} \times \frac{1\ \text{cm}^3}{1\ \cancel{\text{mL}}} = \boxed{4.6\ \text{cm}^3}$$

Converting the volume from mL to L:

$$\text{Volume (L)} = 4.6\ \cancel{\text{mL}} \times \frac{10^{-3}\ \text{L}}{1\ \cancel{\text{mL}}} = \boxed{4.6\times10^{-3}\ \text{L}}$$

Check The units and magnitudes of the answers seem correct, and it makes sense that the volume in mL would have a number 1000 times larger than the same volume in L.

FOLLOW-UP PROBLEM 1.4 Within a cell, proteins are synthesized on particles called ribosomes. Assuming ribosomes are spherical, what is the volume (in dm^3 and μL) of a ribosome whose average diameter is 21.4 nm (V of a sphere $= \frac{4}{3}\pi r^3$)? Draw a road map to show how you plan the solution. (See Brief Solutions.)

Mass The quantity of matter an object contains is its **mass.** The SI unit of mass is the **kilogram (kg),** the only base unit whose standard is an object—a platinum-iridium cylinder kept in France—and the only one whose name has a prefix.*

The terms *mass* and *weight* have distinct meanings:

- *Mass is constant* because an object's quantity of matter cannot change.
- ***Weight*** *is variable* because it depends on the local gravitational field.

Because the strength of the gravitational field varies with altitude, you (and other objects) weigh slightly less on a high mountain than at sea level.

Does this mean that a sample weighed on laboratory balances in Miami (sea level) and in Denver (about 1.7 km above sea level) give different results? No, because these balances measure mass, not weight. Mechanical balances compare the object's mass with masses built into the balance, so the local gravitational field pulls on them equally. Electronic (analytical) balances generate an electric field that counteracts the local field, and the current needed to restore the pan to zero is converted to the equivalent mass and displayed.

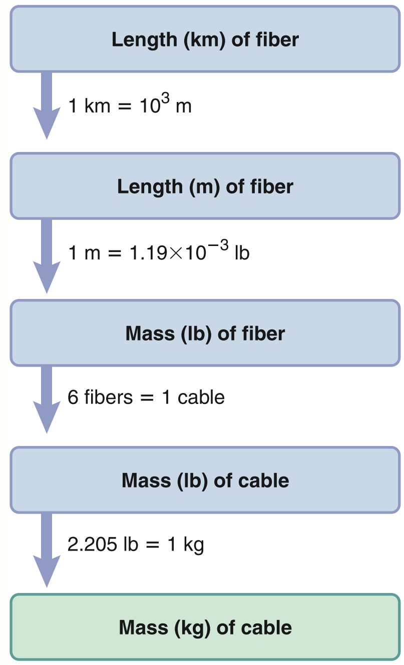
Sample Problem 1.5 Converting Units of Mass

Problem Many international computer communications are carried by optical fibers in cables laid along the ocean floor. If one strand of optical fiber weighs 1.19×10^{-3} lb/m, what is the mass (in kg) of a cable made of six strands of optical fiber, each long enough to link New York and Paris (8.84×10^3 km)?

*The names of the other base units are used as the root words, but for units of mass we attach prefixes to the word "gram," as in "microgram" and "kilogram"; thus, we say "milligram," never "microkilogram."

Plan We have to find the mass of cable (in kg) from the given mass/length of fiber (1.19×10^{-3} lb/m), number of fibers/cable (6), and length of cable (8.84×10^{3} km). Let's first find the mass of one fiber and then the mass of cable. As shown in the road map, we convert the length of one fiber from km to m and then find its mass (in lb) by converting m to lb. Then we multiply the fiber mass by 6 to get the cable mass, and finally convert lb to kg.

Road Map



Solution Converting the fiber length from km to m:

$$\text{Length (m) of fiber} = 8.84\times10^{3}\ \cancel{\text{km}} \times \frac{10^{3}\ \text{m}}{1\ \cancel{\text{km}}} = 8.84\times10^{6}\ \text{m}$$

Converting the length of one fiber to mass (lb):

$$\text{Mass (lb) of fiber} = 8.84\times10^{6}\ \cancel{\text{m}} \times \frac{1.19\times10^{-3}\ \text{lb}}{1\ \cancel{\text{m}}} = 1.05\times10^{4}\ \text{lb}$$

Finding the mass of the cable (lb):

$$\text{Mass (lb) of cable} = \frac{1.05\times10^{4}\ \text{lb}}{1\ \cancel{\text{fiber}}} \times \frac{6\ \cancel{\text{fibers}}}{1\ \text{cable}} = 6.30\times10^{4}\ \text{lb/cable}$$

Converting the mass of the cable from lb to kg:

$$\text{Mass (kg) of cable} = \frac{6.30\times10^{4}\ \cancel{\text{lb}}}{1\ \text{cable}} \times \frac{1\ \text{kg}}{2.205\ \cancel{\text{lb}}} = 2.86\times10^{4}\ \text{kg/cable}$$

Check The units are correct. Let's think through the relative sizes of the answers to see if they make sense: The number of m should be 10^3 larger than the number of km. If 1 m of fiber weighs about 10^{-3} lb, about 10^7 m should weigh about 10^4 lb. The cable mass should be six times as much, or about 6×10^4 lb. Since 1 lb is about $\frac{1}{2}$ kg, the number of kg should be about half the number of lb.

FOLLOW-UP PROBLEM 1.5 An intravenous nutrient solution is delivered to a hospital patient at a rate of 1.5 drops per second. If a drop of solution weighs 65 mg on average, how many kilograms are delivered in 8.0 h? Draw a road map to show how you plan the solution. (See Brief Solutions.)

Density The **density (*d*)** of an object is its mass divided by its volume:

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \tag{1.1}$$

We isolate each of these variables by treating density as a conversion factor:

$$\text{Mass} = \text{volume} \times \text{density} = \cancel{\text{volume}} \times \frac{\text{mass}}{\cancel{\text{volume}}}$$

or,

$$\text{Volume} = \text{mass} \times \frac{1}{\text{density}} = \cancel{\text{mass}} \times \frac{\text{volume}}{\cancel{\text{mass}}}$$

Because volume can change with temperature, so can density. But, at a given temperature and pressure, the *density of a substance is a characteristic physical property and, thus, has a specific value*.

The SI unit of density is kilograms per cubic meter (kg/m^3), but in chemistry, density has units of g/L (g/dm^3) or g/mL (g/cm^3) (Table 1.4). Note that the densities of gases are much lower than those of liquids or solids (also see Figure 1.2).

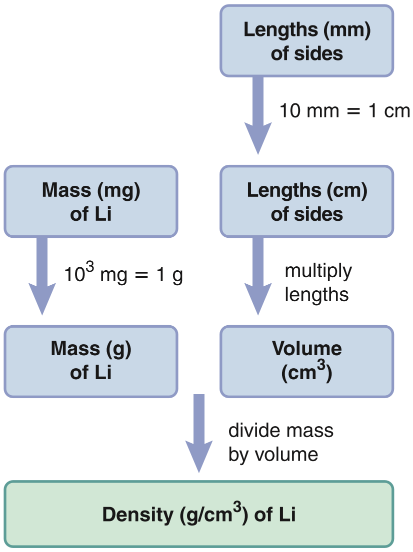
Table 1.4 Densities of Some Common Substances*

Substance	Physical State	Density (g/cm^3)
Hydrogen	gas	0.0000899
Oxygen	gas	0.00133
Grain alcohol	liquid	0.789
Water	liquid	0.998
Table salt	solid	2.16
Aluminum	solid	2.70
Lead	solid	11.3
Gold	solid	19.3

*At room temperature (20°C) and normal atmospheric pressure (1 atm).

Sample Problem 1.6 Calculating Density from Mass and Volume

Road Map



Problem Lithium, a soft, gray solid with the lowest density of any metal, is a key component of advanced batteries, such as the one in your laptop. A slab of lithium weighs 1.49×10^3 mg and has sides that are 20.9 mm by 11.1 mm by 11.9 mm. Find the density of lithium in g/cm^3.

Plan To find the density in g/cm^3, we need the mass of lithium in g and the volume in cm^3. The mass is 1.49×10^3 mg, so we convert mg to g. We convert the lengths of the three sides from mm to cm, and then multiply them to find the volume in cm^3. Dividing the mass by the volume gives the density (see the road map).

Solution Converting the mass from mg to g:

$$\text{Mass (g) of lithium} = 1.49 \times 10^3 \, \cancel{\text{mg}} \left(\frac{10^{-3}\text{ g}}{1\, \cancel{\text{mg}}} \right) = 1.49 \text{ g}$$

Converting side lengths from mm to cm:

$$\text{Length (cm) of one side} = 20.9\, \cancel{\text{mm}} \times \frac{1 \text{ cm}}{10\, \cancel{\text{mm}}} = 2.09 \text{ cm}$$

Similarly, the other side lengths are 1.11 cm and 1.19 cm.
Multiplying the sides to get the volume:

$$\text{Volume (cm}^3) = 2.09 \text{ cm} \times 1.11 \text{ cm} \times 1.19 \text{ cm} = 2.76 \text{ cm}^3$$

Calculating the density:

$$\text{Density of lithium} = \frac{\text{mass}}{\text{volume}} = \frac{1.49 \text{ g}}{2.76 \text{ cm}^3} = \boxed{0.540 \text{ g/cm}^3}$$

Check Since 1 cm = 10 mm, the number of cm in each length should be $\frac{1}{10}$ the number of mm. The units for density are correct, and the size of the answer (~0.5 g/cm^3) seems correct since the number of g (1.49) is about half the number of cm^3 (2.76). Also, the problem states that lithium has a very low density, so this answer makes sense.

FOLLOW-UP PROBLEM 1.6 The piece of galena in Sample Problem 1.4 has a volume of 4.6 cm^3. If the density of galena is 7.5 g/cm^3, what is the mass (in kg) of that piece of galena? Draw a road map to show how you plan the solution. (See Brief Solutions.)

Temperature There is a key distinction between temperature and heat:

- **Temperature (*T*)** is *a measure* of how hot or cold one object is relative to another.
- **Heat** is *the energy* that flows from the object with the higher temperature to the object with the lower temperature. When you hold an ice cube, it feels like the "cold" flows into your hand, but actually, heat flows from your hand to the ice.

In the laboratory, we measure temperature with a **thermometer,** a narrow tube containing a fluid that expands when heated. When the thermometer is immersed in a substance hotter than itself, heat flows from the substance through the glass into the fluid, which expands and rises in the thermometer tube. If a substance is colder than the thermometer, heat flows to the substance from the fluid, which contracts and falls within the tube.

We'll consider three temperature scales: the Celsius (°C, formerly called centigrade); the Kelvin (K), which is preferred in scientific work (although the Celsius scale is still used frequently); and the Fahrenheit (°F) scales. The SI base unit of temperature is the **kelvin** (**K,** with no degree sign, °). In the United States, the Fahrenheit scale is used for weather reporting, body temperature, and so forth.

The three scales differ in the size of the unit and/or the temperature of the zero point. Figure 1.6 shows the freezing and boiling points of water in the three scales. The **Celsius scale** sets water's freezing point at 0°C and its boiling point (at normal atmospheric pressure) at 100°C. The **Kelvin (absolute) scale** uses the *same size degree* as the Celsius scale—$\frac{1}{100}$ of the difference between the freezing and boiling points of water—but it has a *different zero point;* that is, 0 K, or *absolute zero,* equals −273.15°C. Thus, in the Kelvin scale, *all temperatures are positive;* for example, water freezes at +273.15 K (0°C) and boils at +373.15 K (100°C).

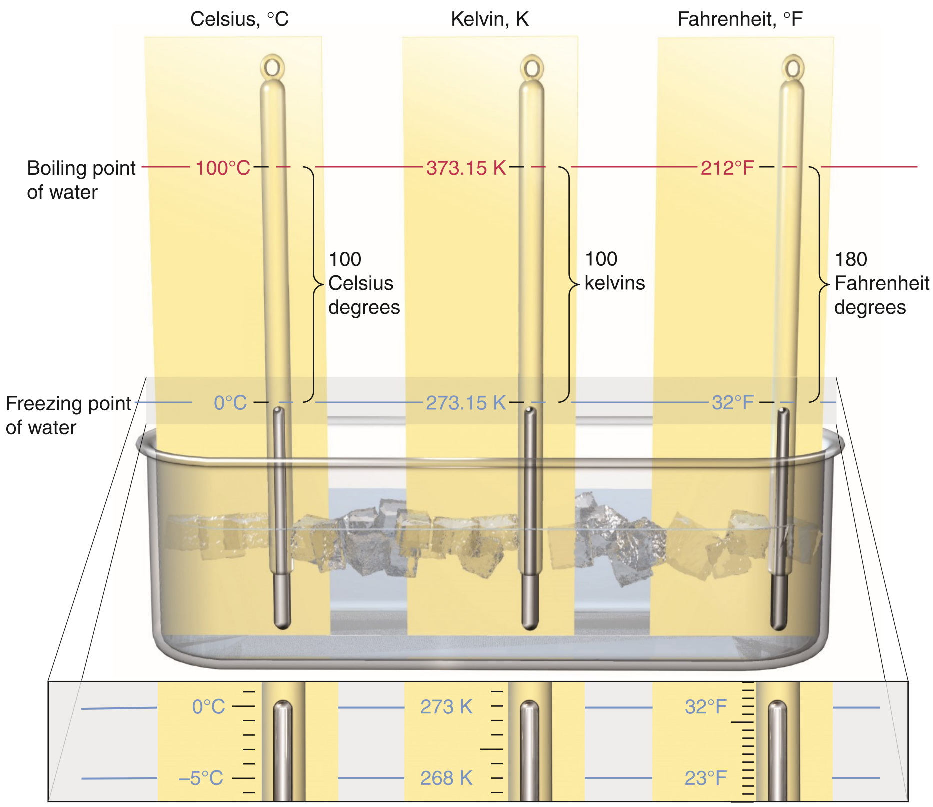


Figure 1.6 **Freezing and boiling points of water in the Celsius, Kelvin (absolute), and Fahrenheit scales.** At the bottom of the figure, a portion of each of the three thermometer scales is expanded to show the sizes of the units.

We convert between the Celsius and Kelvin scales by remembering the different zero points: 0°C = 273.15 K, so

$$T\ (\text{in K}) = T\ (\text{in °C}) + 273.15 \quad \textbf{(1.2)}$$

And, therefore,

$$T\ (\text{in °C}) = T\ (\text{in K}) - 273.15 \quad \textbf{(1.3)}$$

The Fahrenheit scale differs from the other scales in its zero point *and* in the size of its degree. Water freezes at 32°F and boils at 212°F. Therefore, 180 Fahrenheit degrees (212°F − 32°F) represents the same temperature change as 100 Celsius degrees (or 100 kelvins). Because 100 Celsius degrees equal 180 Fahrenheit degrees,

$$1 \text{ Celsius degree} = \tfrac{180}{100} \text{ Fahrenheit degrees} = \tfrac{9}{5} \text{ Fahrenheit degrees}$$

To convert a temperature from °C to °F, first change the degree size and then adjust the zero point:

$$T\ (\text{in °F}) = \tfrac{9}{5}T\ (\text{in °C}) + 32 \quad \textbf{(1.4)}$$

To convert a temperature from °F to °C, do the two steps in the opposite order: adjust the zero point and then change the degree size. In other words, solve Equation 1.4 for T (in °C):

$$T\ (\text{in °C}) = [T\ (\text{in °F}) - 32]\tfrac{5}{9} \quad \textbf{(1.5)}$$

Table 1.5 compares the three temperature scales.

Table 1.5 The Three Temperature Scales

Scale	Unit	Size of Degree (Relative to K)	Freezing Point of H_2O	Boiling Point of H_2O	T at Absolute Zero	Conversion
Kelvin (absolute)	kelvin (K)	—	273.15 K	373.15 K	0 K	to °C (Equation 1.2)
Celsius	Celsius degree (°C)	1	0°C	100°C	−273.15°C	to K (Equation 1.3) to °F (Equation 1.4)
Fahrenheit	Fahrenheit degree (°F)	$\frac{5}{9}$	32°F	212°F	−459.67°F	to °C (Equation 1.5)

Sample Problem 1.7 Converting Units of Temperature

Problem A child has a body temperature of 38.7°C, and normal body temperature is 98.6°F. Does the child have a fever? What is the child's temperature in kelvins?

Plan To see if the child has a fever, we convert from °C to °F (Equation 1.4) and compare it with 98.6°F. Then we convert the temperature in °C to K (Equation 1.2).

Solution Converting the temperature from °C to °F:

$$T\text{ (in °F)} = \tfrac{9}{5}T\text{ (in °C)} + 32 = \tfrac{9}{5}(38.7\text{°C}) + 32 = 101.7\text{°F};$$

Converting the temperature from °C to K:

$$T\text{ (in K)} = T\text{ (in °C)} + 273.15 = 38.7\text{°C} + 273.15 = 311.8\text{ K}$$

Check From everyday experience, you know that 101.7°F is a reasonable temperature for someone with a fever. In the second step, we can check for a large error as follows: 38.7°C is almost 40°C, and 40 + 273 = 313, which is close to our answer.

FOLLOW-UP PROBLEM 1.7 Mercury melts at 234 K, lower than any other pure metal. What is its melting point in °C and °F?

Time The SI base unit of time is the **second (s),** which is now based on an atomic standard. The most recent version of the atomic clock is accurate to within 1 second in 20 million years! The atomic clock measures the oscillations of microwave radiation absorbed by gaseous cesium atoms cooled to around 10^{-6} K: 1 second is defined as 9,192,631,770 of these oscillations. Chemists now use lasers to measure the speed of extremely fast reactions that occur in a few picoseconds (10^{-12} s) or femtoseconds (10^{-15} s).

Extensive and Intensive Properties

Some variables are *dependent* on the amount of substance present; these are called **extensive properties.** On the other hand, **intensive properties** are *independent* of the amount of substance. Mass and volume, for example, are extensive properties, but density is an intensive property. Thus, a gallon of water has four times the mass of a quart of water, but it also has four times the volume, so the density, the *ratio* of mass to volume, is the same for both samples.

Another important example concerns heat, an extensive property, and temperature, an intensive property: a vat of boiling water has more heat, that is, more energy, than a cup of boiling water, but both samples have the same temperature.

Summary of Section 1.4

- The SI unit system consists of seven base units and numerous derived units.
- Exponential notation and prefixes based on powers of 10 are used to express very small and very large numbers.
- The SI base unit of length is the meter (m); on the atomic scale, the nanometer (nm) and picometer (pm) are used commonly.
- Volume (V) units are derived from length units, and the most important volume units are the cubic meter (m^3) and the liter (L).
- The *mass* of an object—the quantity of matter in it—is constant. The SI unit of mass is the kilogram (kg). The *weight* of an object varies with the gravitational field.
- Density (d) is a characteristic physical property of a substance and is the ratio of its mass to its volume.
- Temperature (T) is a measure of the relative hotness of an object. Heat is energy that flows from an object at higher T to one at lower T.
- Temperature scales differ in the size of the degree unit and/or the zero point. For scientific uses, temperature is measured in kelvins (K) or degrees Celsius (°C).
- Extensive properties, such as mass, volume, and energy, depend on the amount of a substance. Intensive properties, such as density and temperature, do not.

1.5 • UNCERTAINTY IN MEASUREMENT: SIGNIFICANT FIGURES

All measuring devices—balances, pipets, thermometers, and so forth—are made to limited specifications, and we use our imperfect senses and skills to read them. Therefore, we can *never* measure a quantity exactly; put another way, every measurement includes some **uncertainty.** The device we choose depends on how much uncertainty is acceptable. When you buy potatoes, a supermarket scale that measures in 0.1-kg increments is acceptable; it tells you that the mass is, for example, 2.0 ± 0.1 kg. The "± 0.1 kg" term expresses the uncertainty: the potatoes weigh between 1.9 and 2.1 kg. Needing more certainty than that to weigh a substance, a chemist uses a balance that measures in 0.001-kg increments and finds the substance weighs 2.036 ± 0.001 kg, that is, between 2.035 and 2.037 kg. The greater number of digits in this measurement means we know the mass of the substance with *more certainty* than we know the mass of potatoes.

We *always estimate the rightmost digit* of a measurement. The uncertainty can be expressed with the ± sign, but generally we drop the sign and *assume an uncertainty of one unit in the rightmost digit.* The digits we record, both the certain and the uncertain ones, are called **significant figures.** There are four significant figures in 2.036 kg and two in 2.0 kg. *The greater the number of significant figures, the greater is the certainty of a measurement.* Figure 1.7 shows this point for two thermometers.

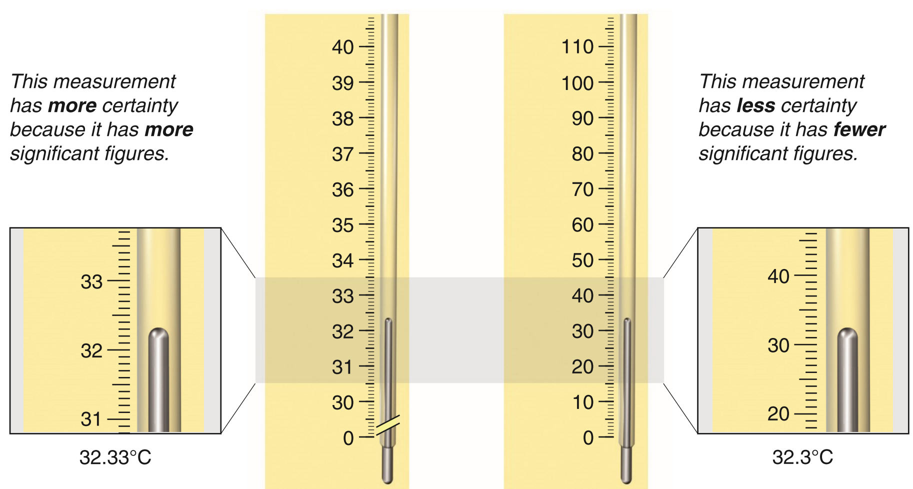


Figure 1.7 The number of significant figures in a measurement. The thermometer on the left is graduated in increments of 0.1°C; the one on the right is graduated in increments of 1°C.

Determining Which Digits Are Significant

When you take a measurement or use one in a calculation, you must know the number of digits that are significant: *all digits are significant, except zeros used only to position the decimal point.* The following procedure applies this point:

1. Make sure the measurement has a decimal point.
2. Start at the left, and move right until you reach the first nonzero digit.
3. Count that digit and every digit to its right as significant.

A complication can arise when zeros end a number:

- If there *is* a decimal point and the zeros lie either after or before it, they *are* significant: 1.1300 g has five significant figures and 6500. has four.
- If there is *no* decimal point, assume the zeros are *not* significant, unless exponential notation clarifies the quantity: 5300 L is *assumed* to have two significant figures, but 5.300×10^3 L has four, 5.30×10^3 L has three, and 5.3×10^3 L has two.
- A terminal decimal point means the zeros before it *are* significant: 500 mL has one significant figure, but 500. mL has three (as do 5.00×10^2 mL and 0.500 L).

Sample Problem 1.8 Determining the Number of Significant Figures

Problem For each of the following quantities, underline the zeros that are significant figures (sf) and determine the total number of significant figures. For (d) to (f), express each quantity in exponential notation first.

(a) 0.0030 L **(b)** 0.1044 g **(c)** 53,069 mL
(d) 0.00004715 m **(e)** 57,600. s **(f)** 0.0000007160 cm^3

Plan We determine the number of significant figures by counting digits, as just described, paying particular attention to the position of zeros in relation to the decimal point, and underline the zeros that are significant.

Solution **(a)** 0.0030 L has 2 sf
(b) 0.1044 g has 4 sf
(c) 53,069 mL has 5 sf
(d) 0.00004715 m, or 4.715×10^{-5} m, has 4 sf
(e) 57,600. s, or 5.7600×10^{4} s, has 5 sf
(f) 0.0000007160 cm^3, or 7.160×10^{-7} cm^3, has 4 sf

Check Be sure that every zero counted as significant comes *after* nonzero digit(s) in the number.

FOLLOW-UP PROBLEM 1.8 For each of the following quantities, underline the zeros that are significant figures and determine the total number of significant figures (sf). For (d) to (f), express each quantity in exponential notation first.

(a) 31.070 mg **(b)** 0.06060 g **(c)** 850.°C
(d) 200.0 mL **(e)** 0.0000039 m **(f)** 0.000401 L

Significant Figures: Calculations and Rounding Off

Measuring several quantities typically results in data with differing numbers of significant figures. In a calculation, we keep track of the number in each quantity so that we don't have more significant figures (more certainty) in the answer than in the data. If we do have too many significant figures, we must **round off** the answer.

The general rule for rounding is that *the least certain measurement sets the limit on certainty for the entire calculation and determines the number of significant figures in the final answer.* Suppose you want to find the density of a new ceramic. You measure the mass of a piece of it on a precise laboratory balance and obtain 3.8056 g; you measure the volume as 2.5 mL by displacement of water in a graduated cylinder. The mass has five significant figures, but the volume has only two. Should you report the density as 3.8056 g/2.5 mL = 1.5222 g/mL or as 1.5 g/mL? The answer with five significant figures implies more certainty than the answer with two. But you didn't measure the volume to five significant figures, so you can't possibly know the density with that much certainty. Therefore, you report 1.5 g/mL, the answer with two significant figures.

Rules for Arithmetic Operations The two rules in arithmetic calculations are

1. *For multiplication and division.* The answer contains the same number of significant figures as there are in the measurement with the *fewest significant figures*. Suppose you want to find the volume of a sheet of a new graphite composite. The length (9.2 cm) and width (6.8 cm) are obtained with a ruler, and the thickness (0.3744 cm) with a set of calipers. The calculation is

$$\text{Volume (cm}^3) = 9.2\text{ cm} \times 6.8\text{ cm} \times 0.3744\text{ cm} = 23.4225\text{ cm}^3 = 23\text{ cm}^3$$

Even though your calculator may show 23.4225 cm^3, you report 23 cm^3, the answer with two significant figures, the same as in the measurements with the lower number of significant figures. After all, if the length and width have two significant figures, you can't possibly know the volume with more certainty.

2. *For addition and subtraction.* The answer has the same number of decimal places as there are in the measurement with the *fewest decimal places*. Suppose you want the total volume after adding water to a protein solution: you have 83.5 mL of solution in a graduated cylinder and add 23.28 mL of water from a buret. The calculation is shown

in the margin. Here the calculator shows 106.78 mL, but you report the volume as 106.8 mL, because the measurement with fewer decimal places (83.5 mL) has one decimal place.

$$\begin{array}{r} 83.5\ \text{mL} \\ +\ 23.28\ \text{mL} \\ \hline 106.78\ \text{mL} \end{array}$$

Answer: Volume = 106.8 mL

Rules for Rounding Off You usually need to round off the final answer to the proper number of significant figures or decimal places. Notice that in calculating the volume of the graphite composite above, we removed the extra digits, but in calculating the total volume of the protein solution, we removed the extra digit and increased the last digit by one. The general rule for rounding is that *the least certain measurement sets the limit on the certainty of the final answer.* Here are detailed rules for rounding off:

1. If the digit removed is *more than 5,* the preceding number increases by 1: 5.379 rounds to 5.38 if you need three significant figures and to 5.4 if you need two.
2. If the digit removed is *less than 5,* the preceding number remains the same: 0.2413 rounds to 0.241 if you need three significant figures and to 0.24 if you need two.
3. If the digit removed *is 5,* the preceding number increases by 1 if it is odd and remains the same if it is even: 17.75 rounds to 17.8, but 17.65 rounds to 17.6. If the 5 is followed only by zeros, rule 3 is followed; if the 5 is followed by nonzeros, rule 1 is followed: 17.6500 rounds to 17.6, but 17.6513 rounds to 17.7.
4. *Always carry one or two additional significant figures through a multistep calculation and round off the final answer* ***only.*** Don't be concerned if you string together a calculation to check a sample or follow-up problem and find that your answer differs in the last decimal place from the one in the book. To show you the correct number of significant figures in text calculations, *we round off intermediate steps,* and that process may sometimes change the last digit.

Note that, unless you set a limit on your calculator, it gives answers with too many figures and you must round the displayed result.

Significant Figures in the Lab The measuring device you choose determines the number of significant figures you can obtain. Suppose an experiment requires a solution made by dissolving a solid in a liquid. You weigh the solid on an analytical balance and obtain a mass with five significant figures. It would make sense to measure the liquid with a buret or a pipet, which measures volumes to more significant figures than a graduated cylinder. If you do choose the cylinder, you would have to round off more digits, and some certainty in the mass value would be wasted (Figure 1.8). With experience, you'll choose a measuring device based on the number of significant figures you need in the final answer.

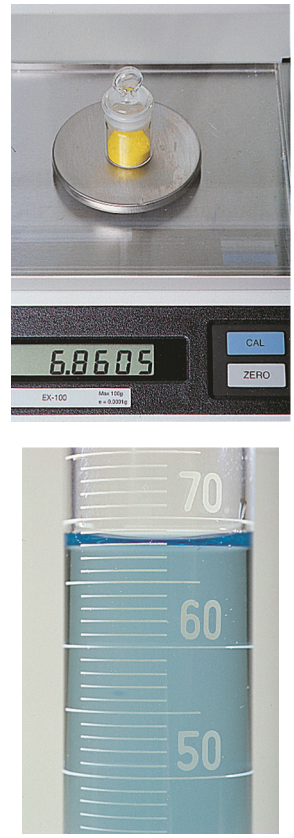


Figure 1.8 Significant figures and measuring devices. The mass measurement (6.8605 g) has more significant figures than the volume measurement (68.2 mL).

Exact Numbers **Exact numbers** have no uncertainty associated with them. Some are part of a unit conversion: by definition, there are exactly 60 minutes in 1 hour, 1000 micrograms in 1 milligram, and 2.54 centimeters in 1 inch. Other exact numbers result from actually counting items: there are exactly 3 coins in my hand, 26 letters in the English alphabet, and so forth. Therefore, unlike a measured quantity, *exact numbers do not limit the number of significant figures in a calculation.*

Sample Problem 1.9 Significant Figures and Rounding

Problem Perform the following calculations and round the answers to the correct number of significant figures:

(a) $\dfrac{16.3521\ \text{cm}^2 - 1.448\ \text{cm}^2}{7.085\ \text{cm}}$ **(b)** $\dfrac{(4.80\times10^4\ \text{mg})\left(\dfrac{1\ \text{g}}{1000\ \text{mg}}\right)}{11.55\ \text{cm}^3}$

Plan We use the rules just presented in the text: **(a)** We subtract before we divide. **(b)** We note that the unit conversion involves an exact number.

Solution (a) $\dfrac{16.3521\ \text{cm}^2 - 1.448\ \text{cm}^2}{7.085\ \text{cm}} = \dfrac{14.904\ \text{cm}^{\cancel{2}}}{7.085\ \cancel{\text{cm}}} = \boxed{2.104\ \text{cm}}$

(b) $\dfrac{(4.80\times10^4\ \cancel{\text{mg}})\left(\dfrac{1\ \text{g}}{1000\ \cancel{\text{mg}}}\right)}{11.55\ \text{cm}^3} = \dfrac{48.0\ \text{g}}{11.55\ \text{cm}^3} = \boxed{4.16\ \text{g/cm}^3}$

Check Note that in (a) we lose a decimal place in the numerator, and in (b) we retain 3 sf in the answer because there are 3 sf in 4.80. Rounding to the nearest whole number is always a good way to check: **(a)** $(16 - 1)/7 \approx 2$; **(b)** $(5\times10^4/1\times10^3)/12 \approx 4$.

FOLLOW-UP PROBLEM 1.9 Perform the following calculation and round the answer to the correct number of significant figures: $\dfrac{25.65\ \text{mL} + 37.4\ \text{mL}}{73.55\ \text{s}\left(\dfrac{1\ \text{min}}{60\ \text{s}}\right)}$

Precision, Accuracy, and Instrument Calibration

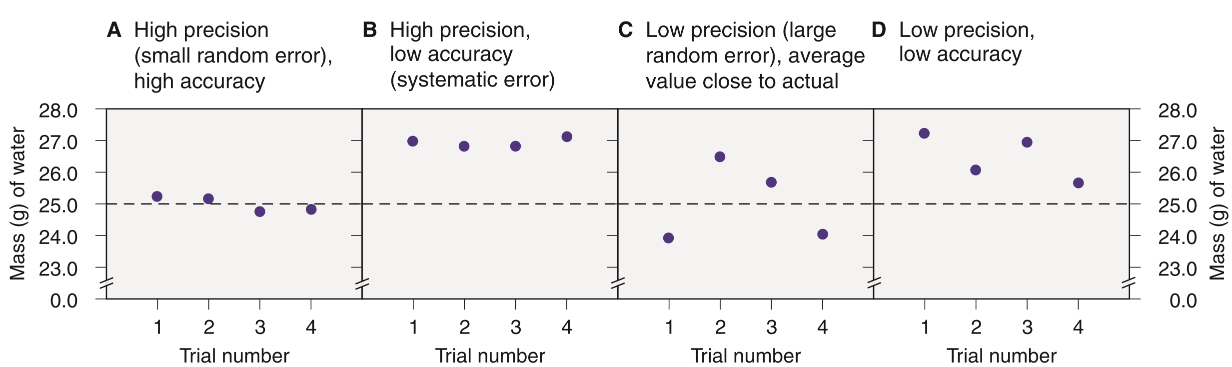
We may use the words "precision" and "accuracy" interchangeably in everyday speech, but for scientific measurements they have distinct meanings. **Precision,** or *reproducibility,* refers to how close the measurements in a series are to each other, and **accuracy** refers to how close each measurement is to the actual value. These terms are related to two widespread types of error:

1. **Systematic error** produces values that are *either* all higher or all lower than the actual value. This type of error is part of the experimental system, often caused by a faulty device or by making the same mistake when taking a reading.
2. **Random error,** in the absence of systematic error, produces values that are higher *and* lower than the actual value. Random error *always* occurs, but its size depends on the measurer's skill and the instrument's precision.

Precise measurements have low random error, that is, small deviations from the average. *Accurate measurements have low systematic error and, generally, low random error.* In some cases, when many measurements have a high random error, the *average* may still be accurate.

Suppose each of four students measures 25.0 mL of water in a preweighed graduated cylinder and then weighs the water *plus* cylinder on a balance. If the density of water is 1.00 g/mL at the temperature of the experiment, the *actual* mass of 25.0 mL of water is 25.0 g. Each student performs the operation four times, subtracts the mass of the empty cylinder, and obtains one of four graphs (Figure 1.9). In graphs A and B, random error is small; that is, precision is high (the weighings are reproducible). In A, however, the accuracy is high as well (all the values are close to 25.0 g), whereas in B the accuracy is low (there is a systematic error). In graphs C and D, random error is large; that is, precision is low. Note, however, that in D there is also a systematic error (all the values are high), whereas in C the average of the values is close to the actual value.

Figure 1.9 Precision and accuracy in a laboratory calibration.



Systematic error can be taken into account through **calibration,** comparing the measuring device with a known standard. The systematic error in graph B, for example, might be caused by a poorly manufactured cylinder that reads "25.0" when it actually contains about 27 mL. If that cylinder had been calibrated, the student could have adjusted all volumes measured with it. The students also should calibrate the balance with standardized masses.

Summary of Section 1.5

- The final digit of a measurement is always estimated. Thus, all measurements have some uncertainty, which is expressed by the number of significant figures.
- The certainty of a calculated result depends on the certainty of the data, so the answer has as many significant figures as in the least certain measurement.
- Excess digits are rounded off in the final answer with a set of rules.
- The choice of laboratory device depends on the certainty needed.
- Exact numbers have as many significant figures as the calculation requires.
- Precision refers to how close values are to each other, and accuracy refers to how close values are to the actual value.
- Systematic errors give values that are either all higher or all lower than the actual value. Random errors give some values that are higher and some that are lower than the actual value.
- Precise measurements have low random error; accurate measurements have low systematic error and low random error.
- A systematic error is often caused by faulty equipment and can be compensated for by calibration.

CHAPTER REVIEW GUIDE

The following sections provide many aids to help you study this chapter. (Numbers in parentheses refer to pages, unless noted otherwise.)

Learning Objectives

These are concepts and skills to review after studying this chapter.

Related section (§), sample problem (SP), and upcoming end-of-chapter problem (EP) numbers are listed in parentheses.

1. Distinguish between physical and chemical properties and changes (§1.1) (SPs 1.1, 1.2) (EPs 1.1, 1.3–1.6)
2. Define the features of the states of matter (§1.1) (EP 1.2)
3. Understand the nature of potential energy and kinetic energy and their interconversion (§1.1) (EPs 1.7–1.8)
4. Understand the scientific approach to studying phenomena and distinguish between observation, hypothesis, experiment, and model (§1.2) (EPs 1.9–1.12)
5. Use conversion factors in calculations (§1.3) (SP 1.3) (EPs 1.13–1.15)
6. Distinguish between mass and weight, heat and temperature, and intensive and extensive properties (§1.4) (EPs 1.16, 1.17, 1.19)
7. Use numerical prefixes and common units of length, mass, volume, and temperature in unit-conversion calculations (§1.4) (SPs 1.4–1.7) (EPs 1.21–1.34, 1.36–1.40)
8. Understand scientific notation and the meaning of uncertainty; determine the number of significant figures and the number of digits after rounding (§1.5) (SPs 1.8, 1.9) (EPs 1.42–1.56)
9. Distinguish between accuracy and precision and between systematic and random error (§1.5) (EPs 1.57–1.59)

Key Terms

These important terms appear in boldface in the chapter and are defined again in the Glossary.

Section 1.1
chemistry (3)
matter (3)
composition (3)
property (3)
physical property (3)
physical change (3)
chemical property (4)
chemical change (chemical reaction) (4)
state of matter (5)
solid (5)
liquid (5)
gas (5)
energy (7)
potential energy (7)
kinetic energy (7)

Section 1.2
scientific method (9)
observation (9)
data (9)
natural law (9)
hypothesis (9)
experiment (9)
variable (9)
controlled experiment (9)
model (theory) (9)

Section 1.3
conversion factor (10)

Section 1.4
SI unit (14)
base (fundamental) unit (14)
derived unit (14)
meter (m) (14)
volume (V) (15)
cubic meter (m^3) (15)
liter (L) (15)
milliliter (mL) (15)
mass (16)

Key Terms continued

kilogram (kg) (16)
weight (16)
density (d) (17)
temperature (T) (18)
heat (18)
thermometer (18)
kelvin (K) (18)
Celsius scale (18)
Kelvin (absolute) scale (18)
second (s) (20)
extensive property (20)
intensive property (20)

Section 1.5
uncertainty (21)
significant figures (21)
round off (22)
exact number (23)
precision (24)
accuracy (24)
systematic error (24)
random error (24)
calibration (25)

Key Equations and Relationships

Numbered and screened concepts are listed for you to refer to or memorize.

1.1 Calculating density from mass and volume (17):

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

1.2 Converting temperature from °C to K (19):

$$T\ (\text{in K}) = T\ (\text{in °C}) + 273.15$$

1.3 Converting temperature from K to °C (19):

$$T\ (\text{in °C}) = T\ (\text{in K}) - 273.15$$

1.4 Converting temperature from °C to °F (19):

$$T\ (\text{in °F}) = \tfrac{9}{5}T\ (\text{in °C}) + 32$$

1.5 Converting temperature from °F to °C (19):

$$T\ (\text{in °C}) = [T\ (\text{in °F}) - 32]\tfrac{5}{9}$$

BRIEF SOLUTIONS TO FOLLOW-UP PROBLEMS

Compare your own solutions to these calculation steps and answers.

1.1 Chemical. The red-and-blue and separate red particles on the left become paired red and separate blue particles on the right.

1.2 (a) Physical. Solid iodine changes to gaseous iodine.
(b) Chemical. Gasoline burns in air to form different substances.
(c) Chemical. In contact with air, substances in torn skin and blood react to form different substances.

1.3 No. of chairs

$$= 3\ \cancel{\text{bolts}} \times \frac{200\ \cancel{\text{m}^2}}{1\ \cancel{\text{bolt}}} \times \frac{(3.281)^2\cancel{\text{ft}^2}}{1\ \cancel{\text{m}^2}} \times \frac{1\ \text{chair}}{31.5\ \cancel{\text{ft}^2}}$$

$$= 205\ \text{chairs}$$

See Road Map 1.3.

Road Map 1.3

Fabric (bolts)
↓ 200 m^2 = 1 bolt
Fabric (m^2)
↓ $(3.281)^2$ ft^2 = 1 m^2
Fabric (ft^2)
↓ 1 chair = 31.5 ft^2
No. of chairs

Road Map 1.4

Diameter (dm)
↓ diameter = 2r
r (dm)
↓ $V = \frac{4}{3}\pi r^2$
V (dm^3)
↓ 1 dm^3 = 1 L; 1 L = 10^6 µL
V (µL)

1.4 Radius of ribosome (dm) $= \dfrac{21.4\ \cancel{\text{nm}}}{2} \times \dfrac{1\ \text{dm}}{10^8\ \cancel{\text{nm}}}$

$$= 1.07\times10^{-7}\ \text{dm}$$

Volume of ribosome (dm^3) $= \frac{4}{3}\pi r^3 = \frac{4}{3}(3.14)(1.07\times10^{-7}\ \text{dm})^3$

$$= 5.13\times10^{-21}\ \text{dm}^3$$

Volume of ribosome (µL) $= (5.13\times10^{-21}\ \cancel{\text{dm}^3})\left(\dfrac{1\ \text{L}}{1\ \cancel{\text{dm}^3}}\right)\left(\dfrac{10\ \mu\text{L}}{1\ \cancel{\text{L}}}\right)$

$$= 5.13\times10^{-15}\ \mu\text{L}$$

See Road Map 1.4.

1.5 Mass (kg) of solution

$$= 8.0\ \cancel{\text{h}} \times \frac{60\ \cancel{\text{min}}}{1\ \cancel{\text{h}}} \times \frac{60\ \cancel{\text{s}}}{1\ \cancel{\text{min}}} \times \frac{1.5\ \cancel{\text{drops}}}{1\ \cancel{\text{s}}}$$

$$\times \frac{65\ \cancel{\text{mg}}}{1\ \cancel{\text{drop}}} \times \frac{1\ \cancel{\text{g}}}{10^3\ \cancel{\text{mg}}} \times \frac{1\ \text{kg}}{10^3\ \cancel{\text{g}}}$$

$$= 2.8\ \text{kg}$$

See Road Map 1.5.

Road Map 1.5

Time (h)
↓ 1 h = 60 min
Time (min)
↓ 1 min = 60 s
Time (s)
↓ 1 s = 1.5 drops
No. of drops
↓ 1 drop = 65 mg
Mass (mg) of solution
↓ 10^3 mg = 1 g
Mass (g) of solution
↓ 10^3 g = 1 kg
Mass (kg) of solution

Road Map 1.6

V (cm^3)
↓ multiply by density
Mass (g)
↓ 10^3 g = 1 kg
Mass (kg)

1.6 $$\text{Mass (kg) of sample} = 4.6\ \cancel{\text{cm}^3} \times \frac{7.5\ \cancel{\text{g}}}{1\ \cancel{\text{cm}^3}} \times \frac{1\ \text{kg}}{10^3\ \cancel{\text{g}}} = 0.034\ \text{kg}$$

See Road Map 1.6.

1.7 T (in °C) $= 234\ \text{K} - 273.15 = -39°\text{C}$
T (in °F) $= \frac{9}{5}(-39°\text{C}) + 32 = -38°\text{F}$
Answer contains two significant figures (see Section 1.5).

1.8 (a) $31.\underline{070}$ mg, 5 sf (b) $0.06\underline{060}$ g, 4 sf
(c) $85\underline{0}$.°C, 3 sf (d) $2.\underline{000}\times10^2$ mL, 4 sf
(e) 3.9×10^{-6} m, 2 sf (f) $4.\underline{0}1\times10^{-4}$ L, 3 sf

1.9 $$\frac{25.65\ \text{mL} + 37.4\ \text{mL}}{73.55\ \cancel{\text{s}}\left(\dfrac{1\ \text{min}}{60\ \cancel{\text{s}}}\right)} = 51.4\ \text{mL/min}$$

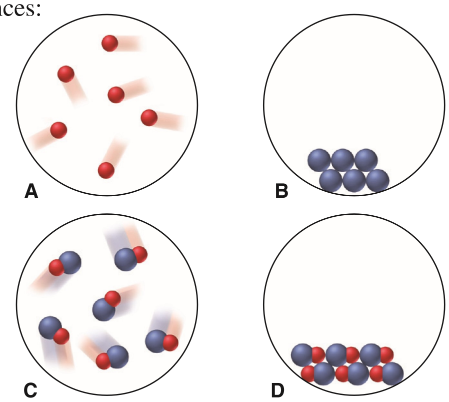
PROBLEMS

Problems with **colored** numbers are answered in Appendix E. Sections match the text and provide the numbers of relevant sample problems. Bracketed problems are grouped in pairs (indicated by a short rule) that cover the same concept. Comprehensive Problems are based on material from any section or previous chapter.

Some Fundamental Definitions

(Sample Problems 1.1 and 1.2)

1.1 Scenes A–D represent atomic-scale views of different samples of substances:



(a) Under one set of conditions, the substances in A and B mix, and the result is depicted in C. Does this represent a chemical or a physical change?
(b) Under a second set of conditions, the same substances mix, and the result is depicted in D. Does this represent a chemical or a physical change?
(c) Under a third set of conditions, the sample depicted in C changes to that in D. Does this represent a chemical or a physical change?
(d) After the change in part (c) has occurred, does the sample have different chemical properties? Physical properties?

1.2 Describe solids, liquids, and gases in terms of how they fill a container. Use your descriptions to identify the physical state (at room temperature) of the following: (a) helium in a toy balloon; (b) mercury in a thermometer; (c) soup in a bowl.

1.3 Define *physical property* and *chemical property.* Identify each type of property in the following statements:
(a) Yellow-green chlorine gas attacks silvery sodium metal to form white crystals of sodium chloride (table salt).
(b) A magnet separates a mixture of black iron shavings and white sand.

1.4 Define *physical change* and *chemical change.* State which type of change occurs in each of the following statements:
(a) Passing an electric current through molten magnesium chloride yields molten magnesium and gaseous chlorine.
(b) The iron in discarded automobiles slowly forms reddish brown, crumbly rust.

1.5 Which of the following is a chemical change? Explain your reasoning: (a) boiling canned soup; (b) toasting a slice of bread; (c) chopping a log; (d) burning a log.

1.6 Which of the following changes can be reversed by changing the temperature: (a) dew condensing on a leaf; (b) an egg turning hard when it is boiled; (c) ice cream melting; (d) a spoonful of batter cooking on a hot griddle?

1.7 For each pair, which has higher potential energy?
(a) The fuel in your car or the gaseous products in its exhaust
(b) Wood in a fire or the ashes after the wood burns

1.8 For each pair, which has higher kinetic energy?
(a) A sled resting at the top of a hill or a sled sliding down the hill
(b) Water above a dam or water falling over the dam

The Scientific Approach: Developing a Model

1.9 How are the key elements of scientific thinking used in the following scenario? While making toast, you notice it fails to pop out of the toaster. Thinking the spring mechanism is stuck, you notice that the bread is unchanged. Assuming you forgot to plug in the toaster, you check and find it is plugged in. When you take the toaster into the dining room and plug it into a different outlet, you find the toaster works. Returning to the kitchen, you turn on the switch for the overhead light and nothing happens.

1.10 Why is a quantitative observation more useful than a nonquantitative one? Which of the following is (are) quantitative? (a) The Sun rises in the east. (b) A person weighs one-sixth as much on the Moon as on Earth. (c) Ice floats on water. (d) A hand pump cannot draw water from a well more than 34 ft deep.

1.11 Describe the essential features of a well-designed experiment.

1.12 Describe the essential features of a scientific model.

Chemical Problem Solving

(Sample Problem 1.3)

1.13 When you convert feet to inches, how do you decide which part of the conversion factor should be in the numerator and which in the denominator?

1.14 Write the conversion factor(s) for
(a) in^2 to m^2 (b) km^2 to cm^2
(c) mi/h to m/s (d) lb/ft^3 to g/cm^3

1.15 Write the conversion factor(s) for
(a) cm/min to in/s (b) m^3 to in^3
(c) m/s^2 to km/h^2 (d) gal/h to L/min

Measurement in Scientific Study

(Sample Problems 1.4 to 1.7)

1.16 Describe the difference between intensive and extensive properties. Which of the following properties are intensive: (a) mass; (b) density; (c) volume; (d) melting point?

1.17 Explain the difference between mass and weight. Why is your weight on the Moon one-sixth that on Earth?

1.18 For each of the following cases, state whether the density of the object increases, decreases, or remains the same:
(a) A sample of chlorine gas is compressed.
(b) A lead weight is carried up a high mountain.
(c) A sample of water is frozen.
(d) An iron bar is cooled.
(e) A diamond is submerged in water.

1.19 Explain the difference between heat and temperature. Does 1 L of water at 65°F have more, less, or the same quantity of energy as 1 L of water at 65°C?

1.20 A one-step conversion is sufficient to convert a temperature in the Celsius scale to the Kelvin scale, but not to the Fahrenheit scale. Explain.

1.21 The average radius of a molecule of lysozyme, an enzyme in tears, is 1430 pm. What is its radius in nanometers (nm)?

1.22 The radius of a barium atom is 2.22×10^{-10} m. What is its radius in angstroms (Å)?

1.23 A small hole in the wing of a space shuttle requires a 20.7-cm^2 patch. (a) What is the patch's area in square kilometers (km^2)? (b) If the patching material costs NASA \$3.25/in^2, what is the cost of the patch?

1.24 The area of a telescope lens is 7903 mm^2. (a) What is the area in square feet (ft^2)? (b) If it takes a technician 45 s to polish 135 mm^2, how long does it take her to polish the entire lens?

1.25 The average density of Earth is 5.52 g/cm^3. What is its density in (a) kg/m^3; (b) lb/ft^3?

1.26 The speed of light in a vacuum is 2.998×10^8 m/s. What is its speed in (a) km/h; (b) mi/min?

1.27 The volume of a certain bacterial cell is 2.56 μm^3. (a) What is its volume in cubic millimeters (mm^3)? (b) What is the volume of 10^5 cells in liters (L)?

1.28 (a) How many cubic meters of milk are in 1 qt (946.4 mL)? (b) How many liters of milk are in 835 gal (1 gal = 4 qt)?

1.29 An empty vial weighs 55.32 g. (a) If the vial weighs 185.56 g when filled with liquid mercury (d = 13.53 g/cm^3), what is its volume? (b) How much would the vial weigh if it were filled with water (d = 0.997 g/cm^3 at 25°C)?

1.30 An empty Erlenmeyer flask weighs 241.3 g. When filled with water (d = 1.00 g/cm^3), the flask and its contents weigh 489.1 g. (a) What is the flask's volume? (b) How much does the flask weigh when filled with chloroform (d = 1.48 g/cm^3)?

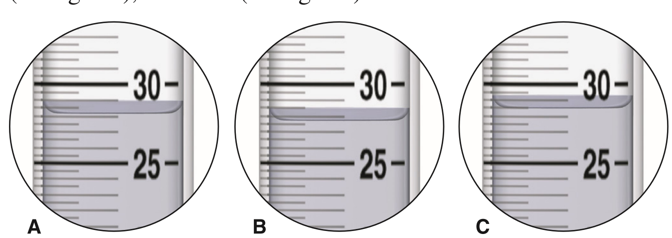
1.31 A small cube of aluminum measures 15.6 mm on a side and weighs 10.25 g. What is the density of aluminum in g/cm^3?

1.32 A steel ball-bearing with a circumference of 32.5 mm weighs 4.20 g. What is the density of the steel in g/cm^3 (V of a sphere = $\frac{4}{3}\pi r^3$; circumference of a circle = $2\pi r$)?

1.33 Perform the following conversions:
(a) 68°F (a pleasant spring day) to °C and K
(b) −164°C (the boiling point of methane, the main component of natural gas) to K and °F
(c) 0 K (absolute zero, theoretically the coldest possible temperature) to °C and °F

1.34 Perform the following conversions:
(a) 106°F (the body temperature of many birds) to K and °C
(b) 3410°C (the melting point of tungsten, the highest for any metallic element) to K and °F
(c) 6.1×10^3 K (the surface temperature of the Sun) to °F and °C

1.35 A 25.0-g sample of each of three unknown metals is added to 25.0 mL of water in graduated cylinders A, B, and C, and the final volumes are depicted in the circles below. Given their densities, identify the metal in each cylinder: zinc (7.14 g/mL), iron (7.87 g/mL), or nickel (8.91 g/mL).



1.36 Anton van Leeuwenhoek, a 17th-century pioneer in the use of the microscope, described the microorganisms he saw as "animalcules" whose length was "25 thousandths of an inch." How long were the animalcules in meters?

1.37 The distance between two adjacent peaks on a wave is called the *wavelength.*
(a) The wavelength of a beam of ultraviolet light is 247 nanometers (nm). What is its wavelength in meters?

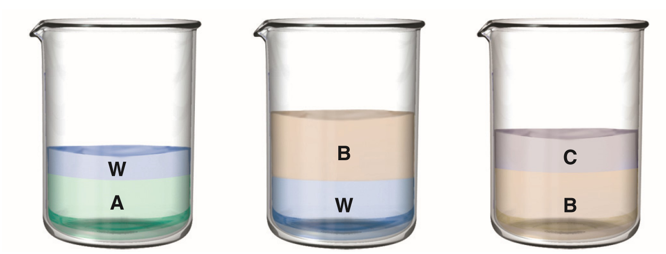
(b) The wavelength of a beam of red light is 6760 pm. What is its wavelength in angstroms (Å)?

1.38 In the early 20th century, thin metal foils were used to study atomic structure. (a) How many in^2 of gold foil with a thickness of 1.6×10^{-5} in could have been made from 2.0 troy oz? (b) If gold cost \$20.00/troy oz at that time, how many cm^2 of gold foil could have been made from \$75.00 worth of gold (1 troy oz = 31.1 g; *d* of gold = 19.3 g/cm^3)?

1.39 A cylindrical tube 9.5 cm high and 0.85 cm in diameter is used to collect blood samples. How many cubic decimeters (dm^3) of blood can it hold (V of a cylinder $= \pi r^2 h$)?

1.40 Copper can be drawn into thin wires. How many meters of 34-gauge wire (diameter $= 6.304\times10^{-3}$ in) can be produced from the copper in 5.01 lb of covellite, an ore of copper that is 66% copper by mass? (*Hint:* Treat the wire as a cylinder: V of cylinder $= \pi r^2 h$; *d* of copper = 8.95 g/cm^3.)

1.41 Each of the beakers depicted below contains two liquids that do not dissolve in each other. Three of the liquids are designated A, B, and C, and water is designated W.



(a) Which of the liquids is (are) more dense than water and which less dense?
(b) If the densities of W, C, and A are 1.0 g/mL, 0.88 g/mL, and 1.4 g/mL, respectively, which of the following densities is possible for liquid B: 0.79 g/mL, 0.86 g/mL, 0.94 g/mL, or 1.2 g/mL?

Uncertainty in Measurement: Significant Figures

(Sample Problems 1.8 and 1.9)

1.42 What is an exact number? How are exact numbers treated differently from other numbers in a calculation?

1.43 All nonzero digits are significant. State a rule that tells which zeros are significant.

1.44 Underline the significant zeros in the following numbers:
(a) 0.41 (b) 0.041 (c) 0.0410 (d) 4.0100×10^4

1.45 Underline the significant zeros in the following numbers:
(a) 5.08 (b) 508 (c) 5.080×10^3 (d) 0.05080

1.46 Carry out the following calculations, making sure that your answer has the correct number of significant figures:

(a) $\dfrac{2.795\text{ m}\times 3.10\text{ m}}{6.48\text{ m}}$

(b) $V = \frac{4}{3}\pi r^3$, where $r = 17.282$ mm

(c) 1.110 cm + 17.3 cm + 108.2 cm + 316 cm

1.47 Carry out the following calculations, making sure that your answer has the correct number of significant figures:

(a) $\dfrac{2.420\text{ g} + 15.6\text{ g}}{4.8\text{ g}}$ (b) $\dfrac{7.87\text{ mL}}{16.1\text{ mL} - 8.44\text{ mL}}$

(c) $V = \pi r^2 h$, where $r = 6.23$ cm and $h = 4.630$ cm

1.48 Write the following numbers in scientific notation:
(a) 131,000.0 (b) 0.00047 (c) 210,006 (d) 2160.5

1.49 Write the following numbers in scientific notation:
(a) 282.0 (b) 0.0380 (c) 4270.8 (d) 58,200.9

1.50 Write the following numbers in standard notation. Use a terminal decimal point when needed:
(a) 5.55×10^3 (b) 1.0070×10^4 (c) 8.85×10^{-7} (d) 3.004×10^{-3}

1.51 Write the following numbers in standard notation. Use a terminal decimal point when needed:
(a) 6.500×10^3 (b) 3.46×10^{-5} (c) 7.5×10^2 (d) 1.8856×10^2

1.52 Carry out each calculation, paying special attention to significant figures, rounding, and units (J = joule, the SI unit of energy; mol = mole, the SI unit for amount of substance):

(a) $\dfrac{(6.626\times10^{-34}\text{ J·s})(2.9979\times10^{8}\text{ m/s})}{489\times10^{-9}\text{ m}}$

(b) $\dfrac{(6.022\times10^{23}\text{ molecules/mol})(1.23\times10^{2}\text{ g})}{46.07\text{ g/mol}}$

(c) $(6.022\times10^{23}\text{ atoms/mol})(2.18\times10^{-18}\text{ J/atom})\left(\dfrac{1}{2^2}-\dfrac{1}{3^2}\right)$,

where the numbers 2 and 3 in the last term are exact

1.53 Carry out each calculation, paying special attention to significant figures, rounding, and units:

(a) $\dfrac{4.32\times10^{7}\text{ g}}{\frac{4}{3}(3.1416)(1.95\times10^{2}\text{ cm})^3}$ (The term $\frac{4}{3}$ is exact.)

(b) $\dfrac{(1.84\times10^{2}\text{ g})(44.7\text{ m/s})^2}{2}$ (The term 2 is exact.)

(c) $\dfrac{(1.07\times10^{-4}\text{ mol/L})^2(3.8\times10^{-3}\text{ mol/L})}{(8.35\times10^{-5}\text{ mol/L})(1.48\times10^{-2}\text{ mol/L})^3}$

1.54 Which statements include exact numbers?
(a) Angel Falls is 3212 ft high.
(b) There are 8 known planets in the Solar System.
(c) There are 453.59 g in 1 lb.
(d) There are 1000 mm in 1 m.

1.55 Which of the following include exact numbers?
(a) The speed of light in a vacuum is a physical constant; to six significant figures, it is 2.99792×10^8 m/s.
(b) The density of mercury at 25°C is 13.53 g/mL.
(c) There are 3600 s in 1 h.
(d) In 2010, the United States had 50 states.

1.56 How long is the metal strip shown below? Be sure to answer with the correct number of significant figures.

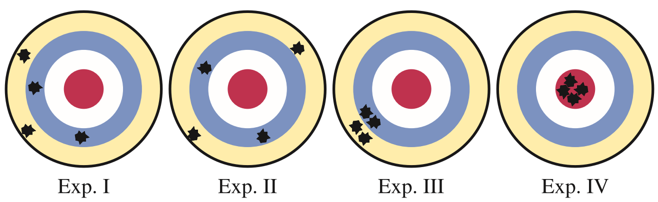
1.57 These organic solvents are used to clean compact discs:

Solvent	Density (g/mL) at 20°C
Chloroform	1.492
Diethyl ether	0.714
Ethanol	0.789
Isopropanol	0.785
Toluene	0.867

(a) If a 15.00-mL sample of CD cleaner weighs 11.775 g at 20°C, which solvent does the sample most likely contain?
(b) The chemist analyzing the cleaner calibrates her equipment and finds that the pipet is accurate to ±0.02 mL, and the balance is accurate to ±0.003 g. Is this equipment precise enough to distinguish between ethanol and isopropanol?

1.58 A laboratory instructor gives a sample of amino-acid powder to each of four students, I, II, III, and IV, and they weigh the samples. The true value is 8.72 g. Their results for three trials are
I: 8.72 g, 8.74 g, 8.70 g II: 8.56 g, 8.77 g, 8.83 g
III: 8.50 g, 8.48 g, 8.51 g IV: 8.41 g, 8.72 g, 8.55 g
(a) Calculate the average mass from each set of data, and tell which set is the most accurate.
(b) Precision is a measure of the average of the deviations of each piece of data from the average value. Which set of data is the most precise? Is this set also the most accurate?
(c) Which set of data is both the most accurate and the most precise?
(d) Which set of data is both the least accurate and the least precise?

1.59 The following dartboards illustrate the types of errors often seen in measurements. The bull's-eye represents the actual value, and the darts represent the data.

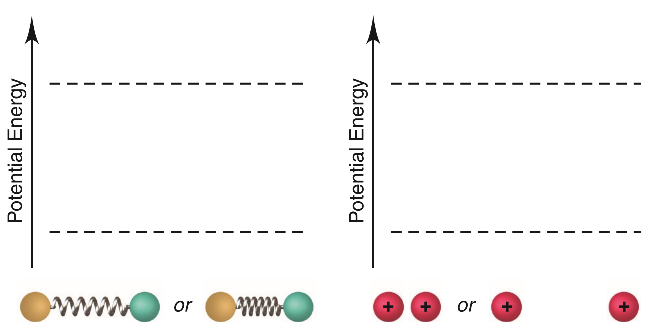


(a) Which experiments yield the same average result?
(b) Which experiment(s) display(s) high precision?
(c) Which experiment(s) display(s) high accuracy?
(d) Which experiment(s) show(s) a systematic error?

Comprehensive Problems

1.60 To make 2.000 gal of a powdered sports drink, a group of students measure out 2.000 gal of water with 500.-mL, 50.-mL, and 5-mL graduated cylinders. Show how they could get closest to 2.000 gal of water, using these cylinders the fewest times.

1.61 Two blank potential energy diagrams appear below. Beneath each diagram are objects to place in the diagram. Draw the objects on the dashed lines to indicate higher or lower potential energy and label each case as more or less stable:

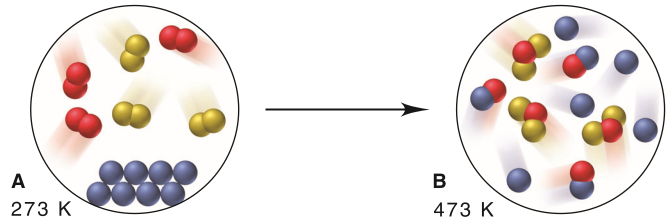


(a) Two balls attached to a relaxed *or* a compressed spring

(b) Two positive charges near *or* apart from each other

1.62 Soft drinks are about as dense as water (1.0 g/cm^3); many common metals, including iron, copper, and silver, have densities around 9.5 g/cm^3. (a) What is the mass of the liquid in a standard 12-oz bottle of diet cola? (b) What is the mass of a dime? (*Hint:* A stack of five dimes has a volume of about 1 cm^3.)

1.63 The scenes below illustrate two different mixtures. When mixture A at 273 K is heated to 473 K, mixture B results.



(a) How many different chemical changes occur?
(b) How many different physical changes occur?

1.64 Suppose your dorm room is 11 ft wide by 12 ft long by 8.5 ft high and has an air conditioner that exchanges air at a rate of 1200 L/min. How long would it take the air conditioner to exchange the air in your room once?

1.65 In 1933, the United States went off the international gold standard, and the price of gold increased from \$20.00 to \$35.00/troy oz. The twenty-dollar gold piece, known as the double eagle, weighed 33.436 g and was 90.0% gold by mass. (a) What was the value of the gold in the double eagle before and after the 1933 price change? (b) How many coins could be made from 50.0 troy oz of gold? (c) How many coins could be made from 2.00 in^3 of gold (1 troy oz = 31.1 g; d of gold = 19.3 g/cm^3)?

1.66 Bromine is used to prepare the pesticide methyl bromide and flame retardants for plastic electronic housings. It is recovered from seawater, underground brines, and the Dead Sea. The average concentrations of bromine in seawater (d = 1.024 g/mL) and the Dead Sea (d = 1.22 g/mL) are 0.065 g/L and 0.50 g/L, respectively. What is the mass ratio of bromine in the Dead Sea to that in seawater?

1.67 An Olympic-size pool is 50.0 m long and 25.0 m wide. (a) How many gallons of water (d = 1.0 g/mL) are needed to fill the pool to an average depth of 4.8 ft? (b) What is the mass (in kg) of water in the pool?

1.68 At room temperature (20°C) and pressure, the density of air is 1.189 g/L. An object will float in air if its density is less than that of air. In a buoyancy experiment with a new plastic, a chemist creates a rigid, thin-walled ball that weighs 0.12 g and has a volume of 560 cm^3.
(a) Will the ball float if it is evacuated?
(b) Will it float if filled with carbon dioxide (d = 1.830 g/L)?
(c) Will it float if filled with hydrogen (d = 0.0899 g/L)?
(d) Will it float if filled with oxygen (d = 1.330 g/L)?
(e) Will it float if filled with nitrogen (d = 1.165 g/L)?
(f) For any case in which the ball will float, how much weight must be added to make it sink?

1.69 Asbestos is a fibrous silicate mineral with remarkably high tensile strength. But it is no longer used because airborne asbestos particles can cause lung cancer. Grunerite, a type of asbestos, has a tensile strength of 3.5×10^2 kg/mm^2 (thus, a strand of grunerite with a 1-mm^2 cross-sectional area can hold up to 3.5×10^2 kg).

The tensile strengths of aluminum and Steel No. 5137 are 2.5×10^4 lb/in^2 and 5.0×10^4 lb/in^2, respectively. Calculate the cross-sectional areas (in mm^2) of wires of aluminum and of Steel No. 5137 that have the same tensile strength as a fiber of grunerite with a cross-sectional area of 1.0 μm^2.

1.70 According to the lore of ancient Greece, Archimedes discovered the displacement method of density determination while bathing and used it to find the composition of the king's crown. If a crown weighing 4 lb 13 oz displaces 186 mL of water, is the crown made of pure gold (d = 19.3 g/cm^3)?

1.71 Earth's oceans have an average depth of 3800 m, a total surface area of 3.63×10^8 km^2, and an average concentration of dissolved gold of 5.8×10^{-9} g/L. (a) How many grams of gold are in the oceans? (b) How many cubic meters of gold are in the oceans? (c) Assuming the price of gold is \$370.00/troy oz, what is the value of gold in the oceans (1 troy oz = 31.1 g; d of gold = 19.3 g/cm^3)?

1.72 For the year 2007, worldwide production of aluminum was 35.1 million metric tons (t). (a) How many pounds of aluminum were produced? (b) What was its volume in cubic feet (1 t = 1000 kg; d of aluminum = 2.70 g/cm^3)?

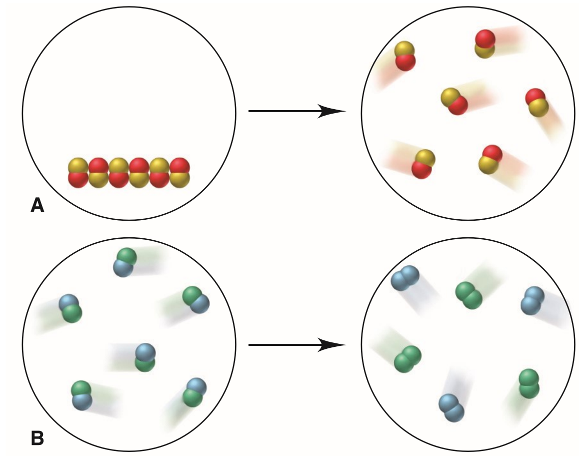
1.73 Liquid nitrogen is obtained from liquefied air and is used industrially to prepare frozen foods. It boils at 77.36 K. (a) What is this temperature in °C? (b) What is this temperature in °F? (c) At the boiling point, the density of the liquid is 809 g/L and that of the gas is 4.566 g/L. How many liters of liquid nitrogen are produced when 895.0 L of nitrogen gas is liquefied at 77.36 K?

1.74 The speed of sound varies according to the material through which it travels. Sound travels at 5.4×10^3 cm/s through rubber and at 1.97×10^4 ft/s through granite. Calculate each of these speeds in m/s.

1.75 If a raindrop weighs 0.52 mg on average and 5.1×10^5 raindrops fall on a lawn every minute, what mass (in kg) of rain falls on the lawn in 1.5 h?

1.76 The Environmental Protection Agency (EPA) proposed a safety standard for microparticulates in air: for particles up to 2.5 μm in diameter, the maximum allowable amount is 50. μg/m^3. If your 10.0 ft × 8.25 ft × 12.5 ft dorm room just meets the EPA standard, how many of these particles are in your room? How many are in each 0.500-L breath you take? (Assume the particles are spheres of diameter 2.5 μm and made primarily of soot, a form of carbon with a density of 2.5 g/cm^3.)

1.77 Scenes A and B depict changes in matter at the atomic scale:



(a) Which show(s) a physical change? (b) Which show(s) a chemical change? (c) Which result(s) in different physical properties? (d) Which result(s) in different chemical properties? (e) Which result(s) in a change in state?

1.78 Earth's surface area is 5.10×10^8 km^2; its crust has a mean thickness of 35 km and a mean density of 2.8 g/cm^3. The two most abundant elements in the crust are oxygen (4.55×10^5 g/t, where t stands for "metric ton"; 1 t = 1000 kg) and silicon (2.72×10^5 g/t), and the two rarest nonradioactive elements are ruthenium and rhodium, each with an abundance of 1×10^{-4} g/t. What is the total mass of each of these elements in Earth's crust?

1.79 The three states of matter differ greatly in their viscosity, a measure of their resistance to flow. Rank the three states from highest to lowest viscosity. Explain in submicroscopic terms.

1.80 If a temperature scale were based on the freezing point (5.5°C) and boiling point (80.1°C) of benzene and the temperature difference between these points was divided into 50 units (called °X), what would be the freezing and boiling points of water in °X? (See Figure 1.6, p. 19.)