A patient arrives at the oncologist's office for his scheduled chemotherapy treatment. The waiting room is completely full, so he suspects that they are running behind schedule. After checking in with the receptionist, he settles in for a long wait. He catches up on his email, sends a few texts to friends, and reads an eight-month-old copy of *Sports Illustrated* from cover to cover before hearing his name called. The nurse leads him from the waiting room, weighs him, takes his temperature, and measures his blood pressure. After preparing his medication, she inserts a needle into a vein in the back of his hand, has another nurse double-check the prescription, and hooks him up to an IV bag. Discovering that all the IV pumps are in use, she decides to set the drip rate manually. After opening the valve and adjusting the flow to 150 milliliters per hour, she makes a note to check for an available pump in 20 minutes, once the patient has received 50 milliliters of IV solution.
In the first chapter of this health science chemistry text, we take a look at the scientific method and at the particular field of science called chemistry. Chemistry is important to health science students because having some knowledge of this field is part of understanding the human body, its diseases, and the medicines used to treat disease. In this chapter we also consider a group of topics related to making measurements.

### CHAPTER 1 OBJECTIVES

After completing this chapter, you should be able to:

1. Explain the terms law, hypothesis, experiment, and theory.
2. Define the terms matter and energy. Describe the three states of matter and the two forms of energy.
3. Describe and give examples of physical properties and physical change.
4. Identify metric, English, and SI units.
5. Express values using scientific notation and metric prefixes.
6. Describe the difference between the terms accurate and precise.
7. Use the correct number of significant figures to report the results of calculations involving measured quantities.
8. Identify conversion factors and use them to convert from one unit to another.
9. Explain the terms density, specific gravity, and specific heat.
10. Recognize the difference between general chemistry, organic chemistry, and biochemistry.
Science is an approach that is used to try to make sense out of how the universe operates, ranging in scale from the very large (understanding how stars form) to the very small (understanding the behavior of the tiny particles from which everything is made). The knowledge gained from scientific studies has impacted our lives in many ways. For example, the discovery of DNA has led to the use of DNA fingerprinting in solving crimes and the development of genetically engineered crops that are better able to deal with pests and pesticides. Studies on energy production are leading to the development of cleaner ways to power our cars, including hydrogen, electricity, and electric/gasoline hybrids. Science has also played an important role in our ever-improving ability to treat diseases. CT and MRI scanners and many of the therapeutic drugs (including anticancer monoclonal antibodies) used today are available as a result of the careful work of scientists (Figure 1.1).

Because science, as a whole, covers such a wide range of subject areas, it is divided into various branches or disciplines. These include chemistry, biology, biochemistry, geology, astronomy, physics, health science, and others. Chemistry, the branch of science involved with the study of matter and its changes, lays an important groundwork for studies in other fields of science, whether they involve designing artificial antibodies to treat disease, testing drinking water for contaminants, or determining the makeup of planets that orbit nearby stars.

When doing science, regardless of the particular discipline, information is gathered and interpreted using the scientific method. Making observations is an important part of this process. One well-known story regarding the importance of observation involves the English scientist Isaac Newton (1642–1727). Reportedly, seeing an apple fall out of a tree led him to formulate the law of gravity, which states that there is an attractive force between any two objects (in this case, between the earth and an apple). This and other scientific laws are statements that describe things that are consistently and reproducibly observed. While a law does not explain why things happen, it can be used to predict what might happen in the future. For example, the law of gravity does not explain why things fall, but it does allow you to predict what will happen if you jump off a ladder.

Finding explanations for observations and laws is a key component of the scientific method. Based on observations or currently known facts, a hypothesis, a tentative explanation (educated guess), can be constructed. Clinicians, for example, make educated guesses when treating patients. If a patient complains of stomach pains, the clinician will ask a few questions and make a few observations before coming up with a hypothesis (diagnosis) as to the nature of the problem. This hypothesis is based on knowledge of symptoms and diseases.

Once a hypothesis has been constructed, it must be tested by doing careful experiments. To test a hypothesis related to a patient’s illness, a clinician might call for a series of medical tests (experiments) to be run. If the test results support the diagnosis, treatment can begin. If they invalidate the diagnosis, the clinician must revise the hypothesis and look for another cause of the illness.

Experiments must be designed so that the observations made are directly related to the question at hand. For example, if a patient has stomach pains, taking an x-ray of his or her big toe will probably not help find the cause of the illness.

If a hypothesis survives repeated testing, it may become a theory—an experimentally tested explanation of an observed behavior. For a theory to survive, it must be consistent with existing experimental evidence, must accurately predict the results of future experiments, and must explain future observations.

Figure 1.2 shows the interconnections of the various parts of the scientific method—making an observation, forming a hypothesis, performing experiments, and creating a theory. Scientists do not necessarily follow these steps in order, nor do they always use
1.1 The Scientific Method

Repeated success over time
Can other scientists repeat the success?

Success

Partial success

Failure

Make an Observation

Develop a Hypothesis

Carry out Experiments

Interpret Findings

Accept Hypothesis

Discard Hypothesis

Revise Hypothesis

Theory

Communicate to Others

In the scientific method, experiments provide the information used to discard, revise, or accept hypotheses.

FIGURE 1.2

all of the steps. It may be that an existing law suggests a new experiment or that a set of published experiments suggests a radically new hypothesis. Creativity is an important part of science; sometimes new theories arise when someone discovers an entirely new way of interpreting experimental results that hundreds of others had looked at before but could not explain. In addition to creativity, a scientist must have sufficient knowledge of the field to be able to interpret experimental results and to evaluate hypotheses and experiments.

SAMPLE PROBLEM 1.1

The scientific method

Suppose that while rearranging your room a few months ago, you moved your favorite plant. Now you notice that the plant is dying. Which of the following explanations are testable hypotheses?

1. The plant is dying because it was moved to a darker location.
2. The plant is dying because it is sad.
3. The plant is dying because it was moved to a warmer location.

STRATEGY

If a statement is a testable hypothesis, you should be able to suggest an experiment to test it.
Improvements in technology also play an important role in science. The fact that theories are based on experimental observations means that as the scientific instruments used to perform experiments improve, theories may have to be changed. In Section 2.1 two theories of the atom, the fundamental particle from which matter is created, are discussed. One of these theories dates back to the early 1800s, when technology was not very advanced and experiments provided much less information than is obtainable today (Figure 1.3). While the earliest theory of the atom accounted for the observations made up until the early 1800s, once better experimental results were obtained, errors were revealed.

Whether scientists study atoms or inherited diseases, theories must be continually reevaluated and, if necessary, revised as new experiments provide additional information. This change is an expected part of science.

**FIGURE 1.3**

*Modifying theories*  Theories are sometimes revised when improved scientific equipment allows better experimental results to be obtained.
Matter and Energy

In the earlier discussion of the various branches of science, chemistry was described as the study of matter and the changes that it undergoes. This leads to the question “what is matter?” In scientific terms, matter is anything that has mass and occupies space. In everyday terms, this definition includes your body, the air that you breathe, this book, and all of the other material around you.

Science and Medicine

The level of glucose (blood sugar) in the body is controlled by a hormone called insulin. Diabetes is the disease that occurs when insulin is not produced in sufficient amounts or when the body is not sensitive to its effects. As science has progressed over the years, so has our understanding of this disease and our ability to treat it. In the mid-1800s, before it was known that high levels of glucose cause the symptoms of diabetes, some physicians recommended that their diabetic patients eat lots of sugar. Others recommended starvation. Scientific studies in the late 1800s and early 1900s led to an understanding of the role that the pancreas plays in glucose metabolism and to the discovery of insulin, which is produced by the pancreas. Insulin was first used in 1922 to treat diabetes in humans. Because the insulin used then was not very pure, patients were given injections—often painful—of up to 2 teaspoons (10 milliliters) at a time. As the science of isolating and purifying insulin improved, dosages dropped to less than one-tenth of that size. Other advances in the treatment of diabetes included the use of oral drugs to control insulin levels (introduced in 1955), the use of genetically engineered human insulin (introduced in 1982) in place of that isolated from cattle and pigs, and the development of new methods for testing blood glucose levels (Figure 1.4).

\[\text{FIGURE 1.4}\]

Glucose testing (a) The urine of diabetics can contain higher than normal amounts of glucose. Many centuries ago, glucose levels were tested by seeing if ants were attracted to a patient’s urine. Sometimes physicians tasted urine to check for sweetness. (b) When these test strips are dipped in a urine sample, the array of colors produced indicates the amount of glucose present. (c) Measuring blood glucose levels is the most accurate way to keep track of diabetes. Blood glucose monitors require just a small drop of blood.

Did You Know

Diabetes mellitus, or diabetes, gets its name from two Greek words related to symptoms of the disease. The word diabetes refers to excessive urination and mellitus to “honey-sweet urine,” which dates to the time when tasting a patient’s urine was part of the diagnosis.
We can describe matter in terms of **physical properties**, those characteristics that can be determined without changing the **chemical composition** of matter (what it is made of). For example, a piece of silver is shiny, conducts electricity, and can be pounded without breaking, while a cube of sugar is white, tastes sweet, can be crushed, and is odorless. The act of measuring these and other physical properties, including melting point (melting temperature), does not change the composition of matter. Silver is still silver and sugar is still sugar.

Matter is typically found in one of three different physical states or phases—as a **solid**, a **liquid**, or a **gas**. From our direct experience we know that

- **Solids** have fixed shapes and volumes.
- **Liquids** have variable shapes and fixed volumes.
- **Gases** have variable shapes and volumes.

Think about what happens if an opened can of paint gets spilled (Figure 1.5). Whether it is standing upright or lying on its side, the can (a solid) has the same shape and occupies the same volume of space. The paint (a liquid) keeps its original 1 gallon volume but changes its shape as it spreads out across the floor. The paint fumes (a gas) quickly change their shape and volume as they spread through the air in the room.

The particular state in which a substance appears depends, in part, on the strength of the interactions between its particles. The term “particles” refers to atoms, molecules, and ions—all three of which we will learn about in future chapters. For now, let’s just think of a water molecule (H₂O) as a particle built from three smaller particles (2 H atoms and 1 O atom).

The particles in a solid are strongly attracted to one another and are held fairly rigidly in one spot. This is true for each of the water molecules that make up ice—they are locked into place through attractions to other water molecules in the solid (Figure 1.6a). The particles in a liquid, including liquid water, are less strongly attracted and are able to slip and slide past one another (Figure 1.6b). The particles that make up a gas are attracted to one another only very weakly, if at all, and are free to move (Figure 1.6c). In later chapters we will take a close look at the types of forces that can attract one particle to another.

Moving between physical states, including from ice to liquid water to steam, is a type of **physical change**, change in which the **chemical composition** of matter is not altered (Figure 1.7). Melting iron, crushing a cube of sugar, and bending a copper wire are also examples of physical change.
Any time that matter is changed in any way, work has been done. This includes the physical changes just mentioned, as well as walking, running, or turning the pages of this book. All of these activities involve energy, which is defined as the ability to do work and to transfer heat.

Energy can be found in two forms, as potential energy (stored energy) or as kinetic energy (the energy of motion). The water sitting behind a dam has potential energy. When the floodgates are opened and the water begins to pour through, potential energy is converted into kinetic energy.

All matter contains energy, so changes in matter (work) and changes in energy (potential or kinetic) are connected to one another. For example, if you drive a car, some of the potential energy of gasoline is converted into the kinetic energy used to move the pistons in the engine (doing work) and some is converted into heat, a form of kinetic energy related to the motion of the particles from which things are made.

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Above, we saw that the strength of the attractions between particles determines, in part, whether a substance is found as a solid, a liquid, or a gas. Heat also plays a role. For example, boiling water to form steam (gaseous water) requires the addition of heat.

Let us take a look at the effect that heat has on the three phases of water: ice, liquid water, and steam. The water molecules in ice are held in place and have a relatively low kinetic energy. If heat is added to water until it melts, liquid water is formed in which the molecules have a greater kinetic energy than in ice. (The higher the temperature of something, the greater the kinetic energy of the particles from which it is made.) Although the water molecules still interact with one another, their increased motion allows them to move around. If heat is added to water until it boils, steam is formed. The even greater kinetic energy allows the water molecules to separate completely from one another and move freely through the container that holds them.

![Figure 1.8](image)

**FIGURE 1.8**

Phase change of water The energy required to convert ice into water is called the heat of fusion. The energy required to convert water into steam is the heat of vaporization.

---

**SAMPLE PROBLEM 1.3**

**Potential versus kinetic energy**

a. You pick up a rubber band and stretch it. What change takes place in the potential energy of the rubber band?

b. You let go of the rubber band and it snaps back to its original shape. What change takes place in the potential energy of the rubber band? What changes take place in its kinetic energy?

c. Is stretching then releasing the rubber band a physical change?

**STRATEGY**

Recall that potential energy is stored energy and that kinetic energy is the energy of motion.

**SOLUTION**

a. The rubber band contains more stored energy, so its potential energy increases.

b. Its potential energy decreases as it releases back to its original shape. As it snaps, the rubber band’s kinetic energy (motion) initially increases but then decreases.

c. Yes, nothing new is created.

**PRACTICE PROBLEM 1.3**

a. Which has greater potential energy, a cup of coffee held at waist level or one held at shoulder level?

b. Which has greater kinetic energy, a cup of hot coffee or a cup of cold coffee?

---

The units used to measure temperature, including degrees Celsius (°C) and degrees Fahrenheit (°F) are discussed in Section 1.3.
gradually adding heat energy to warm it, we will see an increase in temperature. When the
temperature reaches 0°C (32°F), the melting point of ice or the freezing point of water,
the temperature remains constant—even as more heat is added—until all of the ice has
melted. The energy put in during this melting process is called the **heat of fusion**. With
the continued addition of heat energy, water temperature rises until it reaches 100°C
(212°F), the boiling point of water. As the water begins boiling, the temperature remains
constant as heat is added, until all of the water has been converted to steam. The energy
that goes into converting water from the liquid to the gas phase is called the **heat of
vaporization**. Once the water has all boiled, the addition of more heat causes the tem-
perature of the steam to rise.

This process can be reversed. As heat energy is removed from steam, its temperature
drops. At a temperature of 100°C, where steam condenses to form liquid water, the tem-
perature remains constant until only water is present. Further loss of heat energy lowers
the temperature of water until, at 0°C, water begins to freeze. Again, the temperature
remains at 0°C until all of the water has been converted into ice. Removal of more heat
energy lowers the temperature of the ice.

Under certain conditions, some substances will skip the liquid phase and jump directly
between the liquid and gas phases. The conversion of a solid directly into a gas is called
**sublimation** and the reverse of this process is called **deposition**. Dry ice, solid carbon
dioxide, is a common example of a substance that undergoes sublimation (Figure 1.9).

---

**SAMPLE PROBLEM 1.4**

**Energy and changes in physical state**

It was once common to reduce a fever by applying isopropyl alcohol to the skin. As the
alcohol evaporates (liquid becomes gas), the skin cools. Explain the changes in heat energy
as this process takes place. Note: Reducing a fever this way is no longer recommended.

**STRATEGY**

To answer this question you must decide whether heat energy must be put into or removed
from rubbing alcohol to convert it into a gas.

**SOLUTION**

The heat energy required to convert rubbing alcohol from a liquid to a gas is provided by
the heat in the skin. As heat moves from the skin into the rubbing alcohol, the skin cools.

---

**PRACTICE PROBLEM 1.4**

The boiling point of water is 100°C and that of ethyl alcohol is 78°C. In which liquid are
the particles (molecules) held to one another more strongly?
Measurements consist of two parts: a number and a unit. Saying that you swam for 3 is not very informative—was it 3 minutes, 3 hours, or 3 miles? The number must be accompanied by a unit, a quantity that is used as a standard of measurement (of time, of length, of volume, etc.). The metric system is the measurement system used most often worldwide. In this text we will use metric units and the English units used in the United States (Table 1.1). Occasionally, units of the SI system (an international system of measurement related to the metric system) will be introduced. Table 1.2 lists some of the additional units that are commonly used in medical applications.

**Mass**

Mass is a measure of the amount of matter in a sample—the more matter that it contains, the greater its mass. Units commonly used to measure mass are kilogram (kg), gram (g), and pound (lb). One kilogram is defined as the mass of a standard bar of platinum-iridium alloy (a mixture of the two metals) maintained by the International Bureau of Weights and Measures. One kilogram is equal to 1000 g and 2.205 lb (Figure 1.11a).

The terms “mass” and “weight” are often used interchangeably, but they do not mean exactly the same thing. While mass is related to the amount of matter in an object, weight

<table>
<thead>
<tr>
<th>TABLE 1.1 MEASUREMENT UNITS</th>
<th>TABLE 1.2 SOME MEASUREMENT UNITS USED IN MEDICINE</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Quantity</strong></td>
<td><strong>English Unit</strong></td>
</tr>
<tr>
<td>Mass</td>
<td>Pound (lb)</td>
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*The prefixes “micro” and “milli” are explained in Section 1.4.
is related to the force with which gravity attracts the object. If you weighed 150 pounds on the earth, you would weigh only 25 pounds on the moon, where gravity is 16.5% as strong. While your weight is different on the earth and the moon, your mass is not because you contain the same amount of matter. This distinction between mass and weight is not significant for our purposes, because we will only be dealing with measurements made on earth.

**Length**

In both the metric and SI unit systems, the unit used for measuring length is the meter (m). One meter is defined as the distance that light travels in a vacuum in 1/299,792,458 of a second. Because a meter is equal to 3.281 ft (39.37 in), a meter stick is slightly longer than a yard stick (Figure 1.11 b).

**Volume**

Volume is the amount of space occupied by an object. The standard SI unit of volume measurement is the cubic meter (m$^3$). This unit is equivalent to the space occupied by a cube that is 1 meter on a side: length (1 m) × width (1 m) × height (1 m) = volume (1 m$^3$).

Because one cubic meter is a large volume (equivalent to about 260 gallons), volume is often expressed using other, smaller, units. The metric unit of volume is the liter (L), which is one-thousandth the size of a cubic meter. One quart equals 0.946 L (Figure 1.11 c).

**Temperature**

In the metric system the Celsius (°C) scale is used to measure temperature. On this scale, water freezes at 0°C and boils at 100°C. On the Fahrenheit (°F) scale used in the United States, water freezes at 32°F and boils at 212°F (Figure 1.12). Besides having different numerical values for the freezing and boiling points of water, these two temperature scales have degrees of different sizes. On the Fahrenheit scale there are 180 degrees between the temperatures where water boils and freezes (212°F − 32°F = 180°F). On the Celsius scale, however, there are only 100 degrees over this same range (100°C − 0°C = 100°C). This means that the boiling to freezing range for water has almost twice as many Fahrenheit degrees as Celsius degrees (180/100 = 1.8).

---

**Did You Know?**

In June 2010, the amount of oil leaking into the Gulf of Mexico each day from BP’s Deep Water Horizon oil spill was estimated to be between 20,000 and 100,000 barrels (1 barrel = 42 gallons). At a rate of 50,000 barrels per day, this is enough oil to fill approximately 13 million of Starbucks’ Venti-sized cups or 3.2 Olympic-sized swimming pools.
Scientists often measure temperature using the SI unit called the kelvin (K). A temperature of 0 K, known as absolute zero, is the temperature at which all heat energy has been removed from a sample. On the Kelvin temperature scale, the difference between the freezing point (273.15 K) and the boiling point (373.15 K) of water is 100 degrees, the same as that for the Celsius scale. Notice that the degree symbol (°) is used when expressing temperature in Celsius and Fahrenheit, but not in Kelvin.

**Energy**

In the metric and English measurement systems, the unit for energy is the calorie (cal). One calorie is defined as the amount of energy required to raise the temperature of 1 g of water by 1°C. The SI energy unit, the joule (J), is approximately equal to the energy expended by a human heart each time that it beats. It takes a little more than four joules to equal one calorie (1 cal = 4.184 J).

When you hear the word “calorie,” it might bring food to mind. One food Calorie (Cal) is equal to 1000 cal, which means that an 80 Cal cookie contains 80,000 cal of potential energy.

**SAMPLE PROBLEM 1.5**

**Comparing units**

Which is larger?

a. 1 m or 1 yd
b. 1 lb or 1 g
c. 1 L or 1 m³
d. 1 cal or 1 J

**STRATEGY**

Solving this problem should not require any calculations. Refer to the “Relationships” column in Table 1.1 to get a feel for the relative sizes of these units.

**SOLUTION**

a. 1 m
b. 1 lb
c. 1 m³
d. 1 cal

**PRACTICE PROBLEM 1.5**

Without doing any calculations, decide which of each pair is the warmer temperature.

a. 273°C or 273 K
b. 32°F or 32°C
c. 0°F or 0 K
1.4 **Scientific Notation, SI and Metric Prefixes**

**Scientific Notation**

When making measurements, particularly in the sciences, there are many times when you must deal with very large or very small numbers. For example, a typical red blood cell has a diameter of about 0.0000075 m. In *scientific notation* (exponential notation) this diameter is written \(7.5 \times 10^{-6}\) m. Values expressed in scientific notation are written as a number between 1 and 10 multiplied by a power of ten. The superscripted number to the right of the ten is called an exponent.

\[
7.5 \times 10^{-6}
\]

An exponent with a positive value tells you how many times to multiply a number by 10,

\[
3.5 \times 10^4 = 3.5 \times 10 \times 10 \times 10 \times 10 = 35000
\]

\[
6.22 \times 10^2 = 6.22 \times 10 \times 10 = 622
\]

while an exponent with a negative value tells you how many times to divide a number by 10.

\[
3.5 \times 10^{-4} = \frac{3.5}{10 \times 10 \times 10 \times 10} = 0.00035
\]

\[
6.22 \times 10^{-2} = \frac{6.22}{10 \times 10} = 0.0622
\]

An easy way to convert a number into scientific notation is to shift the decimal point. For a number that is equal to or greater than 10, shift the decimal point to the left until you get a number between 1 and 10. The number of spaces that you moved the decimal place is the new exponent (see Table 1.3).

\[
35000 = 3.5 \times 10^4
\]

\[
285.2 = 2.852 \times 10^2
\]

\[
8300000 = 8.3 \times 10^6
\]

For a number smaller than 1, shift the decimal point to the right until you get a number between 1 and 10. Put a negative sign in front of the number of spaces that you moved the decimal place and make this the new exponent.

\[
0.00035 = 3.5 \times 10^{-4}
\]

\[
0.0445 = 4.45 \times 10^{-2}
\]

\[
0.0000003554 = 3.554 \times 10^{-8}
\]

A benefit of using scientific notation is that it allows you to compare very large or small numbers without having to count zeros. A particular virus, for example, is 0.00000010 m in diameter, while a human hair has a width of about 0.00010 m (Figure 1.13). How much smaller is this virus than a human hair? In scientific notation, the virus diameter (\(1.0 \times 10^{-7}\) m) is about 3 powers of ten (\(10^3 = 1000\)) times smaller in diameter than the hair (\(1.0 \times 10^{-8}\) m).

---

**Table 1.3 Scientific Notation**

<table>
<thead>
<tr>
<th>Number</th>
<th>Scientific Notation</th>
<th>Exponent</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.0001</td>
<td>(1 \times 10^{-4})</td>
<td>−4</td>
</tr>
<tr>
<td>0.001</td>
<td>(1 \times 10^{-3})</td>
<td>−3</td>
</tr>
<tr>
<td>0.01</td>
<td>(1 \times 10^{-2})</td>
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<tr>
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<td>3</td>
</tr>
<tr>
<td>10000</td>
<td>(1 \times 10^4)</td>
<td>4</td>
</tr>
</tbody>
</table>

**Figure 1.13**

*Ahuman hair* A human hair has a thickness of about 0.00010 m (\(1.0 \times 10^{-4}\) m).
While we may be more accustomed to expressing large numbers in ordinary notation, scientific notation is often used. For example, the human body contains somewhere on the order of 30,000,000,000,000 red blood cells. In scientific notation, this number is expressed as $3 \times 10^{13}$.

**SAMPLE PROBLEM 1.6**

**Using scientific notation**

Convert each number into scientific notation.

a. 0.0144  
   b. 144  
   c. 36.32  
   d. 0.0000098

**STRATEGY**

The decimal point is shifted to the left for numbers equal to or greater than 10 and shifted to the right for numbers smaller than 1.

**SOLUTION**

a. $1.44 \times 10^{-2}$  
   b. $1.44 \times 10^{2}$  
   c. $3.632 \times 10^{1}$  
   d. $9.8 \times 10^{-6}$

**PRACTICE PROBLEM 1.6**

One-millionth of a liter of blood contains about 5 million red blood cells. Express this volume of blood and this number of cells using scientific notation.

**SI and Metric Prefixes**

When making measurements, scientific notation is not the only way to express large and small numbers. Another approach that can be used is to create larger and smaller units by attaching a prefix that indicates how the new unit relates to the original (see Table 1.4).

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Symbol</th>
<th>Multiplier</th>
</tr>
</thead>
<tbody>
<tr>
<td>giga</td>
<td>G</td>
<td>$10^9$</td>
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<tr>
<td>mega</td>
<td>M</td>
<td>$10^6$</td>
</tr>
<tr>
<td>kilo</td>
<td>k</td>
<td>$10^3$</td>
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<tr>
<td>hecto</td>
<td>h</td>
<td>$10^2$</td>
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<tr>
<td>deka</td>
<td>da</td>
<td>$10^1$</td>
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<td>1</td>
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<tr>
<td>deci</td>
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<tr>
<td>centi</td>
<td>c</td>
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<tr>
<td>milli</td>
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<tr>
<td>micro</td>
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<tr>
<td>nano</td>
<td>n</td>
<td>$10^{-9}$</td>
</tr>
<tr>
<td>pico</td>
<td>p</td>
<td>$10^{-12}$</td>
</tr>
</tbody>
</table>
For example, drugs are often administered in milliliter (mL) volumes. The prefix “milli” indicates that the original unit, in this case the liter, has been multiplied by $10^{-3}$.

$$1 \text{ milliliter (mL)} = 1 \times 10^{-3} \text{ L}$$

Similarly, distance can be measured in kilometers. The prefix “kilo” indicates that the meter unit of length has been multiplied by $10^3$.

$$1 \text{ kilometer (km)} = 1 \times 10^3 \text{ m}$$

The prefixes in Table 1.4 are most often applied to metric and SI units, so you will encounter units such as centimeters, microliters, and milligrams per deciliter. While still technically correct, it is very unlikely that you will see prefixes applied to English units (milliquarts, kiloinsches, etc.).

Often the best choice for a prefix is one that has a value near that of the number. For example, the prefix “kilo” would be appropriate for a number in the thousands, while the prefix “centi” would work well for one in the hundredths.

$$5500 \text{ meters} = 5.5 \times 10^3 \text{ meters} = 5.5 \text{ kilometers}$$

$$0.032 \text{ meters} = 3.2 \times 10^{-2} \text{ meters} = 3.2 \text{ centimeters}$$

### SAMPLE PROBLEM 1.7

**Using SI and metric prefixes**

a. A small hot tub holds 2000 L of water. Express this volume by adding an appropriate prefix to “liter.”

b. Thirty drops of water corresponds to 0.002 L of water. Express this volume by adding an appropriate prefix to “liter.”

**STRATEGY**

Express each volume using scientific notation and then refer to Table 1.4 to select a prefix.

**SOLUTION**

a. $2000 \text{ L} = 2 \times 10^3 \text{ L} = 2 \text{ kL}$

b. $0.002 \text{ L} = 2 \times 10^{-3} \text{ L} = 2 \text{ mL}$

### PRACTICE PROBLEM 1.7

a. A penny has a diameter of about 0.02 m. Express this distance using an appropriate metric prefix.

b. One thousand cold virus particles placed end to end would span a distance of about 0.000002 m. Express this distance using an appropriate metric prefix.

---

### 1.5 Measurements and Significant Figures

We have just examined some of the units used to report the measured properties of a material. In this section we will address three of the important factors to consider when making measurements: accuracy, precision, and significant figures.

Accuracy is related to how close a measured value is to a true value. Suppose that a patient’s temperature is taken twice and values of 98°F and 102°F are obtained. If the patient’s actual temperature is 103°F, the second measurement is more accurate because it is closer to the true value.
Precision is a measure of reproducibility. The closer that separate measurements come to one another, the more precise they are. Suppose that a patient’s temperature is taken three times and values of 98°F, 99°F, and 97°F are obtained. Another set of temperature measurements gives 90°F, 100°F, and 96°F. The first three measurements are more in agreement with one another, so they are more precise than the second set.

A set of precise measurements is not necessarily accurate and a set of accurate measurements is not necessarily precise. This is illustrated in Figure 1.14, using the game of darts as an example. Figure 1.14(a) shows the results of three shots that are precise, but not accurate—the shots fall close together, but they are not centered on the bull’s-eye. In Figure 1.14(b), the shots are accurate, but not precise, because the shots fall near the bull’s-eye but not close together. Figure 1.14(c) shows three shots that are both accurate and precise.

Significant Figures

The quality of the equipment used to make a measurement is one factor in obtaining accurate and precise results. For example, balances similar to the one shown in Figure 1.15 come in different models. A lower-priced model might report masses to within ±0.1 g, and a higher-priced one to within ±0.001 g.

Suppose that the precision of a balance is such that repeated measurements always agree to within ±0.1 g. On this balance, a U.S. quarter (25 cent coin) might have a reported mass of 5.7 g. This number, 5.7, has two significant figures (those digits in a measurement that are reproducible when the measurement is repeated, plus the first uncertain digit). Here the “7” in 5.7 is uncertain, because the balance reports mass with an error of ±0.1 g. Assuming that the balance is accurate, the actual mass of the quarter may be a little bit more or a little bit less than 5.7 grams.

On a different balance that reports masses with a precision of ±0.001 g, the reported mass of the same quarter might be 5.671 g. Using this measuring device, the mass of the quarter is reported with four significant figures.

For the numbers above (5.7 and 5.671), determining significant figures is straightforward: all of the digits written are significant. Things get a bit trickier when zeros are involved, because zeros that are part of the measurement are significant, while those that only specify the position of the decimal point are not. Table 1.5 summarizes the rules for determining when a digit is significant.

It is important to note that significant figures apply only to measurements, because measurements always contain some degree of error. Numbers have no error when they are obtained by an exact count (there are seven patients sitting in the waiting room) or are defined (12 eggs = 1 dozen, 1 km = 1000 m). These exact numbers have an unlimited number of significant figures.
Measurements and Significant Figures

SAMPLE PROBLEM 1.8

Determining significant figures

Specify the number of significant figures in each measured value.

a. 30.1°C  b. 0.00730 m  c. 7.30 \times 10^3 m  d. 44.50 mL

STRATEGY

All nonzero digits are significant. Zeros, however, are significant only under certain conditions (see Table 1.5).

SOLUTION

a. 3  b. 3  c. 3  d. 4

PRACTICE PROBLEM 1.8

Write each measured value in exponential notation, being sure to give the correct number of significant figures.

a. 7032 cal  b. 88.0 L  c. 0.00005 g  d. 0.06430 lb

<table>
<thead>
<tr>
<th>TABLE 1.5 SIGNIFICANT FIGURES</th>
</tr>
</thead>
<tbody>
<tr>
<td>Scientific notation</td>
</tr>
<tr>
<td>a. All digits, including zeros, are significant (^4)</td>
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<tr>
<td>Ordinary notation</td>
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<tr>
<td>a. All nonzero digits are significant</td>
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<td>b. Zeros placed between nonzero digits are significant</td>
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<tr>
<td>c. Zeros placed at the end of a number when a decimal point is present are significant</td>
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<tr>
<td>d. Zeros placed at the end of a number with no decimal point are not significant</td>
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<tr>
<td>e. Zeros placed at the beginning of a number are not significant</td>
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</tbody>
</table>

\(^4\)When determining significant figures for numbers in scientific notation, the power of 10 is not included.
The percentage of Americans who are obese has risen over the past two decades. According to the available data, in 1990 between 10 and 14% of the U.S. population was obese (Figure 1.16). In 2000, 25% or more of the people in half of the states in the United States were obese. By the year 2010, many states saw obesity percentages rise above 30% of the population. These numbers are of concern to health professionals because being overweight or obese increases a person’s risk of developing health problems. Among the identified overweight- and obesity-related diseases are type II diabetes, heart disease, high blood pressure, osteoarthritis, stroke, sleep apnea, and some cancers.

For years, determining whether someone was overweight involved using height and weight charts. These charts were of limited usefulness because, when it comes to assessing the risk of overweight- and obesity-related disease, body weight is not the main the issue. The primary factor to consider is the percentage of body weight that is due to fat.

Health professionals can determine percentage body fat using a variety of techniques. One of these is the skinfold measurement, in which calipers are used to test the thickness of folds of skin at various places on the body (Figure 1.17). A calculation using the measured values gives body fat percentage. Underwater weighing is another method used to determine body fat levels. Because fat has a lower density than muscle and bone, the more fat that a person has, the less the person will weigh underwater. Once measurements have been made, a set of equations is used to calculate percent body fat. In a different technique called bioelectrical impedance, electrodes are placed on different parts of the body and a low electrical current is applied. Fat is a poorer conductor of electricity than muscle and bone, so the higher the percent body fat, the greater the resistance or impedance to the current.

Although it is not based on direct measurements of percent body fat, Body Mass Index (BMI) is a good predictor of an individual’s risk for overweight- or obesity-related disease. BMI is calculated from a person’s weight and height, using the equation

\[
\text{BMI} = \frac{\text{weight (lb)}}{\text{[height (in)]}^2}
\]

A 5 foot 1 inch (61.0 inch) tall person weighing 145 lb has a BMI of 27.4. According to adult BMI standards (Table 1.6), this person is overweight.

\[
\text{BMI} = 703 \times \frac{\text{weight (lb)}}{\text{[height (in)]}^2} = 703 \times \frac{145}{61.0^2} = 27.4
\]

BMI values are not directly correlated with percent body fat, because age and gender differences exist. On average, at the same BMI, older people have more body fat than younger ones, and women tend to have more body fat than men. In spite of these variations, health professionals can use BMIs to decide if further...
Calculations Involving Significant Figures

Reporting answers with too many or too few significant figures is a problem commonly encountered with calculations involving measured values. The important thing to remember is that calculations should not change the degree of uncertainty in a value.

When doing multiplication or division with measured values, the answer should have the same number of significant figures as the quantity with the fewest. Suppose that you are asked to determine the area of a rectangle. According to your measurements, its width is 5.3 cm and its length is 6.1 cm. Since area = width \times length, you use your calculator to multiply the two, and obtain

\[
\begin{align*}
5.3 \text{ cm} & \quad \text{Two significant figures} \\
\times 6.1 \text{ cm} & \quad \text{Two significant figures} \\
32.33 \text{ cm}^2 & \quad \text{Calculator answer (four significant figures)} \\
\end{align*}
\]

The result given by your calculator has too many significant figures. Each of the original numbers (5.3 and 6.1) has just two significant figures, but the calculator has given an answer with four. Rewriting a number with the proper number of significant figures means that we have to drop the digits that are not significant (in this case, the two to the right of the decimal point) and round off the last digit of the number. We will use the following rules when rounding numbers:

- If the first digit to be removed is 0, 1, 2, 3, or 4, leave the last reported digit unchanged. (57.42 rounds off to 57.4 if three significant figures are needed and to 57 if two significant figures are needed.)
- If the first digit to be removed is 5, 6, 7, 8, or 9, increase the last reported digit by 1. (57.69 rounds off to 57.7 if three significant figures are needed and to 58 if two significant figures are needed.)

The “first digit to be removed” is the digit immediately to the right of the last significant digit in a number.
SAMPLE PROBLEM 1.10

Multiplication and division calculations involving significant figures

Each of the numbers below is measured. Solve the calculations and give the correct number of significant figures.

\[ a. \ 0.12 \times 1.77 \quad b. \ 690.4 \div 12 \quad c. \ 5.444 \times 0.44 \quad d. \ (16.5 \times 0.1140) \div 3.5 \]

**STRATEGY**

These problems all involve multiplication or division, so the answers should have the same number of significant figures as the original quantity with the fewest significant figures.

**SOLUTION**

\[ a. \ 0.21 \quad b. \ 58 \quad c. \ 150 \text{ or } 1.5 \times 10^2 \quad (2 \text{ significant figures}) \quad d. \ 0.54 \]

**PRACTICE PROBLEM 1.10**

Each of the numbers below is measured. Solve the calculations and give the correct number of significant figures.

\[ a. \ 53.4 \times 489.6 \quad b. \ 6.333 \times 10^{-4} \times 5.77 \times 10^3 \quad c. \ (5 \times 989.5) \div 16.3 \quad d. \ (0.45 \times 6) \times 3.14 \]

When doing addition or subtraction with measured values, the answer should have the same number of decimal places as the quantity with the fewest decimal places. Suppose that you are given three mass measurements and are asked to calculate the total mass:

\[
\begin{align*}
13.5 \text{ g} & \quad \text{One decimal place} \\
2.335 \text{ g} & \quad \text{Three decimal places} \\
+ 653 \text{ g} & \quad \text{Zero decimal places} \\
668.835 \text{ g} & \quad \text{Calculator answer (three decimal places)} \\
(669 \text{ g}) & \quad \text{Correct answer (zero decimal places)}
\end{align*}
\]
SAMPLE PROBLEM 1.11

Addition and subtraction calculations involving significant figures

Each of the numbers below is measured. Solve the calculations and give the correct number of significant figures.

a. $4.55 + 1.8$
   
   b. $690.4 - 12.67$
   
   c. $5.44 - 0.444$
   
   d. $16.5 + 0.114 + 3.55$

STRATEGY

These problems all involve addition or subtraction, so the answers should have the same number of decimal places as the original quantity with the fewest number of decimal places.

SOLUTION

a. $6.4$
   
   b. $677.7$
   
   c. $5.00$
   
   d. $20.2$

PRACTICE PROBLEM 1.11

Each of the numbers below is measured. Solve the calculations and give the correct number of significant figures.

a. $53.4 + 489.6$
   
   b. $5.77 	imes 10^3 - 6.333 	imes 10^{-4}$
   
   c. $8.2 + 121 + 16.3$
   
   d. $45.32 - 6 + 6.75$

Body Temperature

When you go in for a medical checkup, a health professional will almost always begin by taking your temperature. This is done because running a fever is a sign of illness. What should your temperature be? A temperature of 98.6°F (37.0°C), measured orally, is considered normal. This normal temperature is actually an average of the typical range of oral body temperatures (97.2–99.9°F) recorded for healthy people.

The human body is divided into two different temperature zones, the core and the shell, so temperature readings will vary depending on which part of your body is measured. The body’s internal core, which holds the organs of the abdomen, chest, and head, is held at a constant temperature. The outer shell, that part of the body nearest the skin, is used to insulate the core. Shell temperatures fluctuate, depending on whether the body is trying to keep or to lose heat, and typically shell temperatures run about 1°F lower than core temperatures.

Rectal temperature measurements are a good way to determine the core body temperature. While oral measurements can indicate core temperature, readings may be incorrect if the thermometer is not placed correctly in the mouth. Hot or cold drinks can also affect the results of oral temperature measurements.

Tympanic membrane (eardrum) measurements give an indication of the core temperature of the brain, while axillary (armpit) and temporal artery (an artery in the head that runs near the temple) give the shell temperature. Like oral measurements, these three techniques are prone to error.

How are temperatures measured?

A variety of methods can be used to take someone’s temperature. The “low tech” method used by countless parents is touch—does your child’s forehead feel hot? As you might expect, this is not the most reliable technique.

For centuries the mercury thermometer has been used to measure temperature. Its operation is based on the fact that mercury expands as it gets warmer—the higher the temperature, the longer the column of mercury in a thermometer. These thermometers, used for rectal, oral, and axillary temperature measurements, have fallen out of favor because they can be difficult to read, can transmit infection when not cleaned properly, and, if broken, can expose people to toxic mercury.

The digital thermometer is one alternative to the mercury thermometer. The operation of this thermometer is based on a thermistor, a device that conducts electricity better the higher the temperature. Digital thermometers, like mercury thermometers, are used to measure rectal, oral, and axillary temperatures. A digital thermometer pacifier has been developed for infant use.

A third type of thermometer measures temperature by detecting infrared (IR) energy, a form of energy that is associated with heat. Tympanic membrane and temporal artery thermometers (Figure 1.18), which operate using IR energy, are quick to use but, in the case of the tympanic thermometers, can give false readings.
CHAPTER 1  Science and Measurements

Because overheating can be a problem for athletes, there has been interest in finding a way to measure core temperature during exercise. A pill-like temperature sensor originally developed by NASA allows trainers and coaches to do just that. The sensor (Figure 1.19), which is swallowed about two hours before exercise to allow it to reach the intestines, emits a signal that can be detected wirelessly when a handset is held to the small of an athlete’s back. It takes about a day and a half for the indigestible sensor to pass completely through the body.

![Remote temperature sensing](image)

**FIGURE 1.19** Remote temperature sensing This indigestible temperature sensor is swallowed. As it moves through the body it wirelessly transmits core temperature readings to a detector.

1.6 CONVERSION FACTORS AND THE FACTOR LABEL METHOD

What is your height in inches and in centimeters? What is the volume of a cup of coffee in milliliters? Answering these questions requires that you convert from one unit into another.

Some unit conversions are simple enough that you can probably do them in your head—six eggs are half a dozen and twenty-four inches are two feet. Solving other conversions may require a systematic approach called the factor label method, which uses conversion factors to transform one unit into another. Conversion factors are derived from the numerical relationship between two units.

Suppose that a 185 lb patient is prescribed a drug whose recommended dosage is listed in terms of kilograms of body weight. To administer the correct dose, you must convert the patient’s pound weight into kilograms. Converting from pounds to kilograms makes use of the equality \(2.205 \text{ lb} = 1 \text{ kg}\) (Table 1.1). Two different conversion factors can be created from this relationship, the first of which is produced by dividing both sides of the equality by 1 kg.

\[
\frac{2.205 \text{ lb}}{1 \text{ kg}} = \frac{1 \text{ kg}}{1 \text{ kg}} = 1 \text{ conversion factor: } \frac{2.205 \text{ lb}}{1 \text{ kg}}
\]

The second conversion factor is created by dividing both sides of the equality by 2.205 lb.

\[
1 \text{ kg} = \frac{2.205 \text{ lb}}{2.205 \text{ lb}} \text{ conversion factor: } \frac{1 \text{ kg}}{2.205 \text{ lb}}
\]

What is the kilogram weight of a 185 lb patient? To answer this question using the factor label method, we multiply 185 lb by the appropriate conversion factor (equal to 1).
In this case the conversion factor to use is the one that has the desired new unit in the numerator. This allows the original units to cancel one another (Figure 1.20).

$$185 \text{ lb} \times \frac{1 \text{ kg}}{2.205 \text{ lb}} = 83.9 \text{ kg}$$

Looking at this answer, you might wonder why it is reported with three significant figures. In the equality $1 \text{ kg} = 2.205 \text{ lb}$, the “1” is an exact number and has an unlimited number of significant figures. The value with the fewest significant figures is 185 lb.

Let us try another one. A vial contains 15 mL of blood serum. Convert this volume into liters. Converting from milliliters into liters uses the equality $1 \text{ mL} = 1 \times 10^{-3} \text{ L}$ (Table 1.4). The two conversion factors derived from this relationship are

$$\frac{1 \text{ mL}}{1 \times 10^{-3} \text{ L}} \quad \text{and} \quad \frac{1 \times 10^{-3} \text{ L}}{1 \text{ mL}}$$

and the conversion factor to use is the one with the new unit (L) in the numerator.

$$15 \text{ mL} \times \frac{1 \times 10^{-3} \text{ L}}{1 \text{ mL}} = 1.5 \times 10^{-2} \text{ L}$$

If you do not have access to a direct relationship between two different units, making a unit conversion may require more than one step. Suppose that you are asked to convert the average volume of blood pumped by one beat of your heart (0.070 L) from liters into cups. What is this volume in cups? You may not know the direct relationship between liters and cups, but $0.946 \text{ L} = 1 \text{ qt}$ (Table 1.1). This gives the conversion factors

$$\frac{1 \text{ qt}}{0.946 \text{ L}} \quad \text{and} \quad \frac{0.946 \text{ L}}{1 \text{ qt}}$$

which means that

$$0.070 \text{ L} \times \frac{1 \text{ qt}}{0.946 \text{ L}} = 0.074 \text{ qt}$$

Knowing that there are 4 cups in one quart gives the equation

$$0.074 \text{ qt} \times \frac{4 \text{ cups}}{1 \text{ qt}} = 0.30 \text{ cup}$$

so 0.070 L is the same volume as 0.30 cup.

The two steps of this conversion can be taken care of at once by incorporating both conversion factors into one equation.

$$0.070 \text{ L} \times \frac{1 \text{ qt}}{0.946 \text{ L}} \times \frac{4 \text{ cups}}{1 \text{ qt}} = 0.30 \text{ cup}$$

---

**FIGURE 1.20**

**Math errors in medicine** A search of the recent literature shows a number of studies related to the potentially harmful effects that math errors can have on patients. In a 2008 study reported in the *Annals of Internal Medicine*, fourteen doctors were told that a 5-year-old patient was having a serious reaction caused by a peanut allergy, and needed immediate treatment with 0.12 mg of epinephrine. Upon being given a bottle containing a 1 mg/mL solution, just eleven of the fourteen doctors calculated the correct dose (0.12 mL).
SAMPLE PROBLEM 1.13

Unit conversions

a. An over-the-counter (nonprescription) cough syrup contains 7.5 mg of dextromethorphan in every 5 mL. The recommended dose of dextromethorphan for a 44 lb child is 10.0 mg. How many milliliters of cough syrup should be given?

b. For a 55 lb child, the recommended dose of dextromethorphan is 12.5 mg. How many milliliters of cough syrup should be given?

STRATEGY

In part a, you are being asked to convert from a 10 mg dose of dextromethorphan to milliliters of cough syrup. For the cough syrup, the relationship between these units (7.5 mg dextromethorphan = 5 mL) can be used to make a conversion factor.

SOLUTION

a. \[10.0 \text{ mg dextromethorphan} \times \frac{5 \text{ mL cough syrup}}{7.5 \text{ mg dextromethorphan}} = 7 \text{ mL cough syrup}\]

b. \[12.5 \text{ mg dextromethorphan} \times \frac{5 \text{ mL cough syrup}}{7.5 \text{ mg dextromethorphan}} = 8 \text{ mL cough syrup}\]

PRACTICE PROBLEM 1.13

The 44 lb child is given a cold tablet that contains 5 mg of dextromethorphan and is then given 5 mL of the cough syrup mentioned in Sample Problem 1.13a. Has the child received greater than the recommended dose?
Temperature Conversions

To convert between degrees Fahrenheit and degrees Celsius, one of the two equations below is used.

\[ ^\circ F = (1.8 \times ^\circ C) + 32 \]
\[ ^\circ C = \frac{^\circ F - 32}{1.8} \]

Between the freezing point (32°F, 0°C) and boiling point (212°F, 100°C) of water there are 180°F and 100°C. The ratio 180/100 equals 1.8, which is the source of this term in the equations—a degree Fahrenheit is 1.8 times smaller than a degree Celsius. The 32 comes from the different freezing point for water on the two temperature scales.

The relationship between degrees Celsius and kelvins is

\[ K = ^\circ C + 273.15 \]
\[ ^\circ C = K - 273.15 \]

These two equations are simpler than the ones used to relate °F and °C because kelvins and Celsius degrees are the same size.

Let us see how a temperature conversion would be carried out. On a warm summer day, the temperature reaches 85°F. What is this temperature in °C?

\[ ^\circ C = \frac{^\circ F - 32}{1.8} = \frac{85 - 32}{1.8} = 29^\circ C \]

**SAMPLE PROBLEM 1.14**

**Temperature conversions**

Liquid nitrogen (N\textsubscript{2}), which has a freezing point of −210°C, is often used to remove warts and to treat precancerous skin lesions. Convert this temperature into kelvins.

**STRATEGY**

To solve this problem you must use an equation that relates °C and K.

**SOLUTION**

\[ K = ^\circ C + 273.15 = -210 + 273.15 = 63 K \]

**PRACTICE PROBLEM 1.14**

The liquid helium used in magnetic resonance imagers (MRIs) has a temperature of 4.1 K. Convert this temperature into °C and °F.

---

**1.7 Density, Specific Gravity, and Specific Heat**

Earlier in the chapter we saw that units related to mass and volume can be used to describe the properties of matter. Another unit, density, is related to both mass and volume. The density of a substance is the amount of mass contained in a given volume. Density is usually expressed in g/cm\textsuperscript{3} for solids, in g/mL for liquids, and in g/L for gases (Table 1.7). At 20°C, 1.00 g of water occupies a volume of 1.00 mL and 1.00 grams of salt has a volume of 0.461 cm\textsuperscript{3}. Their respective densities are 1.00 g/mL and 2.17 g/cm\textsuperscript{3}.

\[
\text{Density of water} = \frac{\text{mass (g)}}{\text{volume (mL)}} = \frac{1.00 \text{ g}}{1.00 \text{ mL}} = 1.00 \text{ g/mL}
\]

\[
\text{Density of salt} = \frac{\text{mass (g)}}{\text{volume (cm}^3\text{)}} = \frac{1.00 \text{ g}}{0.461 \text{ cm}^3} = 2.17 \text{ g/cm}^3
\]
Temperature is specified when the density of a substance is reported because, usually, as temperature varies, so does density. This relationship between temperature and density is the basis for using mercury (\(\text{Hg}\)) in thermometers. At a temperature of 0°C, where the density of \(\text{Hg}\) is 13.60 g/mL, 100 g of \(\text{Hg}\) occupies a volume of 7.35 mL. At 100°C, where the density is 13.35 g/mL, the same 100 g of \(\text{Hg}\) has a volume of 7.49 mL. This increase in volume is what causes the mercury in a thermometer to rise.

Density can be used as a conversion factor that relates mass and volume. For example, at 20°C isopropyl alcohol (rubbing alcohol) has a density of 0.785 g/mL. Another way of saying this is that 0.785 g of isopropyl alcohol occupies a volume of 1.00 mL. One conversion factor based on this relationship is equal to the density.

\[
\frac{0.785 \text{ g}}{1 \text{ mL}} = \text{ conversion factor}
\]

The second conversion factor is formed by dividing both sides of the equality by 0.785 g.

\[
\frac{0.785 \text{ g}}{1 \text{ mL}} = \frac{1 \text{ mL}}{0.785 \text{ g}} = \text{ conversion factor}
\]

Suppose that a lab experiment asks you to measure out 25.0 g of isopropyl alcohol. With liquids it is usually easier to pour a specific volume of liquid than to weigh a particular mass, so one of the two conversion factors above can be used to convert from grams of isopropyl alcohol to milliliters of isopropyl alcohol.

\[
25.0 \text{ g isopropyl alcohol} \times \frac{1 \text{ mL isopropyl alcohol}}{0.785 \text{ g isopropyl alcohol}} = 31.8 \text{ mL isopropyl alcohol}
\]

Similarly, density can be used to convert from volume of a substance into mass of a substance. What is the mass, in grams, of 2.15 L of isopropyl alcohol? This can be calculated by first converting the volume (liters) into the volume unit that appears in the density of the liquid (milliliters). Next, volume is multiplied by the appropriate conversion factor to obtain the mass in grams.

\[
2.15 \text{ L} \times \frac{1 \text{ mL}}{1 \times 10^{-3} \text{ L}} = 2.15 \times 10^3 \text{ mL}
\]

\[
2.15 \times 10^3 \text{ mL} \times \frac{0.785 \text{ g}}{1 \text{ mL}} = 1.69 \times 10^3 \text{ g}
\]

### TABLE 1.7 DENSITY OF COMMON SUBSTANCES

<table>
<thead>
<tr>
<th>Solids (at 20°C)</th>
<th>Density (g/cm³)</th>
<th>Liquids (at 20°C)</th>
<th>Density (g/mL)</th>
<th>Gases a (at 0°C)</th>
<th>Density b (g/L)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cork</td>
<td>0.25</td>
<td>Isopropyl alcohol</td>
<td>0.785</td>
<td>Hydrogen (H₂)</td>
<td>0.0899</td>
</tr>
<tr>
<td>Fat (human)</td>
<td>0.94</td>
<td>Kerosene</td>
<td>0.82</td>
<td>Helium (He)</td>
<td>0.179</td>
</tr>
<tr>
<td>Bone</td>
<td>1.90</td>
<td>Water</td>
<td>1.00</td>
<td>Air</td>
<td>1.29</td>
</tr>
<tr>
<td>Salt</td>
<td>2.17</td>
<td>Whole blood</td>
<td>1.06</td>
<td>Oxygen (O₂)</td>
<td>1.43</td>
</tr>
<tr>
<td>Lead</td>
<td>11.35</td>
<td>Chloroform</td>
<td>1.49</td>
<td>Chlorine (Cl₂)</td>
<td>3.21</td>
</tr>
</tbody>
</table>

a At a pressure of 1 atm.

b Gas density is so low that values are usually reported in g/L, rather than g/mL.
Specific gravity relates the density of a substance to that of water. Here is an example of how specific gravity can be calculated. At 20°C, the density of isopropyl alcohol is 0.785 g/mL and that of water is 1.00 g/mL. Therefore, the specific gravity of isopropyl alcohol is 0.785.

Specific gravity is likely to vary with temperature because a change in temperature affects the density of water and that of the substance being tested, but not necessarily to the same extent.

A substance with a specific gravity of less than 1 will float in water, while one with a specific gravity of greater than 1 will sink. For example, cork has a specific gravity of 0.25, apples a specific gravity of about 0.74, and cast iron a specific gravity of 7.2.

To measure the density of a substance, you must determine its mass and the volume of space that it occupies at a given temperature. Measurements of specific gravity are usually less involved and can be made using refractometers, hydrometers, or test strips (Figure 1.21).

Specific gravity measurements are useful for determining the acid levels in car batteries, the antifreeze levels in car radiators, and the alcohol content in beer and wine. The specific gravity of urine (a mixture of water and waste products excreted by the kidneys) can be used to diagnose kidney problems. Urine with a high specific gravity has too many waste products dissolved in it, which can indicate dehydration or overproduction of antidiuretic hormone (ADH), which regulates the amount of water in blood serum. High ADH levels can be indicative of stress or trauma and often follow major surgery. Urine with a low specific gravity may be an indication of kidney disease, excess fluid intake, or underproduction of ADH.

**SAMPLE PROBLEM 1.15**

**Calculations involving density**

The density of mercury is 13.60 g/mL at 0°C and 13.35 g/mL at 100°C. What volume, in milliliters, does 55.0 g of mercury occupy at each temperature?

**STRATEGY**

Find a density-based conversion factor that will allow you to convert grams of mercury into milliliters of mercury.

**SOLUTION**

The volume of 55.0 g of mercury at 0°C is 4.04 mL.

\[
55.0 \text{ g-mercury} \times \frac{\text{mL mercury}}{13.60 \text{ g-mercury}} = 4.04 \text{ mL mercury}
\]

The volume of 55.0 g of mercury at 100°C is 4.12 mL.

\[
55.0 \text{ g-mercury} \times \frac{\text{mL mercury}}{13.35 \text{ g-mercury}} = 4.12 \text{ mL mercury}
\]

**PRACTICE PROBLEM 1.15**

a. What is the mass of 2.51 \times 10^{-2} L of mercury at 0°C?

b. What is the mass of the same volume of mercury at 100°C?

**FIGURE 1.21**

Measuring specific gravity

(a) Refractometers measure the extent to which a light beam is bent by a liquid. This refraction is related to specific gravity. (b) The level at which the bulb of a hydrometer floats in a liquid is determined by specific gravity.
CHAPTER 1 Science and Measurements

Making Weight

In some sports, including wrestling, boxing, weight lifting, and judo, athletes compete with others who have a similar weight. “Making weight” refers to the custom of rapidly losing weight to become eligible to compete in a particular weight classification (Figure 1.22). This rapid weight reduction typically involves loss of water (one cup of water weighs about 1/2 pound). Rapid water loss techniques include restricting fluid intake and increasing sweat production by sitting in a sauna or exercising in a hot room while wearing a rubber suit. After qualifying for a lower weight division, the idea is to rehydrate during the time that elapses between the weighing in and competition. An athlete doing this will be heavier and, presumably, stronger than his or her competitors.

It is not clear that making weight puts an athlete at a competitive advantage. Dehydration is known to adversely affect endurance, strength, energy, and motivation. Extreme dehydration can result in kidney and heart failure. In 1997, three college wrestlers died while trying to rapidly lose weight through water loss.

In response to these deaths, the NCAA created new rules to discourage rapid weight loss. These rules include banning the use of saunas and rubber suits for this purpose. A certification process that requires athletes to be hydrated before being eligible for weighing was also put in place. Athletes with a urine specific gravity greater than 1.025, which is at the high end of the normal range (1.002–1.0035), are considered to be dehydrated and ineligible for official weigh in.

We just saw that the density of a substance relates mass and volume. A different unit of measurement, called specific heat, relates energy (in calories), mass (in grams), and temperature (in degrees Celsius). Specific heat is defined as the amount of heat energy required to raise the temperature of 1 gram of the substance by 1°C. Water has a specific heat of 1.000 cal/g °C (Table 1.8), which means that 1.000 cal of heat will raise the temperature of 1 g of water by 1°C.

Specific heat of water \[ \frac{1.000 \text{ cal}}{1 \text{ g} \times 1^\circ \text{C}} = 1.000 \text{ cal/g °C} \]

Specific heat can be used to solve various problems related to mass, temperature, and energy. For example, we might wonder exactly how many calories of heat energy are needed to increase the temperature of 2.5 g of water by 5.0°C. We can apply the factor label method, using specific heat as a conversion factor (Section 1.6). The conversion factors for specific heat calculations involving water are

\[
\frac{1.000 \text{ cal}}{\text{g °C}} \quad \text{and} \quad \frac{1.000 \text{ cal}}{\text{g °C}}
\]

Selecting the conversion factor which allows all units but calories to be canceled gives

\[
2.5 \text{ g} \times 5.0^\circ \text{C} \times \frac{1.000 \text{ cal}}{\text{g °C}} = 13 \text{ cal}
\]

This calculation shows that an input of 13 cal of heat energy will raise the temperature of 2.5 g of water by 5.0°C.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific Heat (cal/g °C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>1.000</td>
</tr>
<tr>
<td>Ice(^a)</td>
<td>0.500</td>
</tr>
<tr>
<td>Steam(^b)</td>
<td>0.480</td>
</tr>
<tr>
<td>Isopropyl alcohol</td>
<td>0.612</td>
</tr>
<tr>
<td>Aluminum</td>
<td>0.215</td>
</tr>
<tr>
<td>Copper</td>
<td>0.0924</td>
</tr>
<tr>
<td>Gold</td>
<td>0.0310</td>
</tr>
</tbody>
</table>

\(^a\)From −10°C to 0°C.

\(^b\)At constant pressure.
Measurements in General Chemistry, Organic Chemistry, and Biochemistry

As we saw in Section 1.1, chemistry is the branch of science that is involved with the study of matter and its changes. While there are many subdisciplines of chemistry, three important ones are general chemistry, organic chemistry, and biochemistry. General chemistry, a study of the fundamental principles of chemistry, deals with a wide range of subjects including the structure of atoms and compounds, chemical reactions, and the units of measurements introduced earlier in this chapter. Organic chemistry is a study of the chemistry of carbon. The organic compounds that we will discuss in later chapters are relatively small, but learning about them will help us understand the characteristics of compounds important in biochemistry. Biochemistry is a study of the chemistry of living things and includes topics ranging from what happens when food is digested to how genetic information is passed from one generation to the next. The paragraphs that follow will provide three examples of how these branches of chemistry relate to what goes on around you.

General Chemistry—A Look at Cadmium

Cadmium is a relatively soft metal with a bluish-white color. Some cadmium-containing compounds are highly colored and have been used as red, orange, and yellow paint pigments. Others have been used to prevent the corrosion of steel and to stabilize plastics. Many of today’s electronic devices are powered by nickel-cadmium batteries.
The high toxicity of cadmium has, in recent years, led to a decrease in its use. Health problems associated with exposure to this metal include vomiting, diarrhea, anemia, weakened bones, damage to the lungs, kidneys, nerves, and brain, as well as cancer.

In 2010, several product recalls were related to the presence of high levels of cadmium. One involved “Best Friends” charm bracelets sold in Claire’s stores (Figure 1.23). The Chinese manufacturers of this and other inexpensive jewelry sometimes substitute cadmium for lead because lead use is regulated due to its high toxicity. Concerns about the presence of cadmium in children's jewelry are not necessarily related to skin exposure, but more to the possible ingestion of this toxic metal caused by putting the jewelry in one’s mouth. In the same year, McDonald’s recalled 12 million promotional glasses for the film *Shrek Forever After* when the paint used on them was found to contain potentially harmful levels of cadmium. In 2011, several states passed laws related to the use of cadmium in children’s jewelry. The jewelry industry in the United States has recently agreed to limit its use.

Cadmium is a heavy metal. In chemistry, this term is generally used to indicate that a metal has a density of $8.65 \text{ g/cm}^3$ or higher. The heavy metals include cadmium ($8.65 \text{ g/cm}^3$), lead ($11.35 \text{ g/cm}^3$), and zinc ($7.14 \text{ g/cm}^3$).

**SAMPLE PROBLEM 1.17**

**Heavy metals**

Which occupies a greater volume, 12.4 g of cadmium or 12.4 g of zinc?

**STRATEGY**

Use the density of each metal as a conversion factor.

**SOLUTION**

The zinc occupies a greater volume.

$$12.4 \text{ g cadmium} \times \frac{1 \text{ cm}^3 \text{ cadmium}}{8.65 \text{ g cadmium}} = 1.43 \text{ cm}^3 \text{ cadmium}$$

$$12.4 \text{ g zinc} \times \frac{1 \text{ cm}^3 \text{ zinc}}{7.14 \text{ g zinc}} = 1.74 \text{ cm}^3 \text{ zinc}$$

**PRACTICE PROBLEM 1.17**

Which has a greater mass, 89.5 cm$^3$ of lead or 89.5 cm$^3$ of cadmium?
Organic Chemistry—a Look at Gasoline

At a gas station, the different grades of gasoline are identified by their octane rating (Figure 1.24). This system is based on the organic compounds isooctane and heptane. Pure isooctane, a very good automotive fuel, is assigned a value of 100, while heptane, a very poor fuel, is given a rating of 0. Gasoline is a mixture of a wide variety of related organic compounds obtained from petroleum, and a mixture with an octane rating of 93 is higher grade (closer to the “ideal” isooctane) than is one with a rating of 87.

Among isooctane’s physical properties are its melting point (−107°C), boiling point (98°C), and density (0.69 g/mL at 20°C). Because of its relative melting and boiling points, isooctane is a liquid at everyday temperatures. Knowing the density of this liquid allows us to relate its mass and volume. For example, the volume occupied by 65 g of isooctane at 20°C is 94 mL.

\[
65 \text{ g isooctane} \times \frac{1 \text{ mL isooctane}}{0.69 \text{ g isooctane}} = 94 \text{ mL isooctane}
\]

We can use density to calculate how much heavier a car is with a full tank of gas than with an empty one. The density of gasoline varies slightly with octane rating, so we will use an average value for the density of gasoline (0.73 g/mL at 20°C) and assume a temperature of 20°C for our calculations. If a car’s gas tank contains 15 gallons of gasoline, what is the mass (in pounds) of the gasoline? Because the density of a liquid is expressed in grams and milliliters, it will be necessary at some point to convert between grams and pounds and between milliliters and gallons. To begin with, 15 gallons is equivalent to 5.7 × 10⁴ mL.

\[
15 \text{ gal} \times \frac{4 \text{ qt}}{\text{gal}} \times \frac{0.946 \text{ L}}{\text{qt}} \times \frac{1 \text{ mL}}{1 \times 10^{-3} \text{ L}} = 5.7 \times 10^4 \text{ mL}
\]

Using density as a conversion factor, we can show that 5.7 × 10⁴ mL of gasoline has a mass of 7.8 × 10³ g.

\[
5.7 \times 10^4 \text{ mL gasoline} \times \frac{0.73 \text{ g gasoline}}{\text{mL gasoline}} = 4.2 \times 10^4 \text{ g gasoline}
\]

This mass of gasoline in grams is equivalent to 170 pounds.

\[
4.2 \times 10^4 \text{ g} \times \frac{1 \text{ lb}}{454 \text{ g}} = 93 \text{ lb}
\]
CHAPTER 1  Science and Measurements

Specific gravity is another property of gasoline to consider. Section 1.7 showed us that the specific gravity of a substance is its density divided by that of water at the same temperature. At 20°C, the density of gasoline is 0.73 g/mL and that of water is 1.00 g/mL, so the specific gravity of gasoline at 20°C is 0.73.

\[
\text{Specific gravity} = \frac{\text{density of isooctane}}{\text{density of water}} = \frac{0.73 \text{ g/mL}}{1.00 \text{ g/mL}} = 0.73
\]

Gasoline does not mix with water and, because its specific gravity is less than 1, it floats on water. This characteristic can be helpful if a gasoline spill occurs. Because it floats, gasoline can be contained by booms (Figure 1.25).

Biochemistry—A Look at DNA

The color of your eyes, your blood type, and whether or not you can roll your tongue are among the many inherited traits that are passed between generations. Deoxyribonucleic acid (DNA) carries this hereditary information. DNA is constructed from four different building blocks (nucleotides), each of which has a relatively flat structure that averages about 0.34 nm in thickness and 1.0 nm wide (Figure 1.26a). The building blocks are combined to form chains (Figure 1.26b), which pair up with one other through interactions between the building blocks in different chains (Figure 1.26c). This double-stranded DNA twists to form a helix (Figure 1.26d).

SAMPLE PROBLEM 1.18

A density and temperature calculation involving gasoline

A change in temperature will alter the density of most liquids. A particular blend of gasoline, for example, has a density of 0.76 g/mL at 8°F and a density of 0.71 g/mL at 98°F. If you fill a container with exactly 1.0 gal of this gasoline on a winter day when the temperature is 8°C, how much will the gasoline weigh (in pounds)?

**STRATEGY**

Use density as a conversion factor to convert from volume of gasoline to mass of gasoline.

**SOLUTION**

The density is reported in grams per milliliter, so the volume of 1.0 gal must be converted into milliliters.

\[
1.0 \text{ gal} \times \frac{4 \text{ qt}}{1 \text{ gal}} \times \frac{0.946 \text{ L}}{1 \text{ qt}} \times \frac{1 \text{ mL}}{1 \times 10^{-3} \text{ L}} = 3.8 \times 10^3 \text{ mL}
\]

Using the density at 8°C as a conversion factor gives the mass of 1.0 gal of gasoline.

\[
3.8 \times 10^3 \text{ mL} \times \frac{0.76 \text{ g}}{1 \text{ mL}} = 2.9 \times 10^3 \text{ g}
\]

The relationship between pounds and grams (1 lb = 454 g) is used as a conversion factor to arrive at the final answer.

\[
2.9 \times 10^3 \text{ g} \times \frac{1 \text{ lb}}{454 \text{ g}} = 6.4 \text{ lb}
\]

**PRACTICE PROBLEM 1.18**

a. If you fill the container from Sample Problem 1.18 with exactly 1.0 gal of gasoline on a summer day when the temperature is 98°C, how many pounds will the gasoline weigh?

b. The amount of gasoline delivered by gasoline pumps is adjusted based on the temperature. The “gallon” of gasoline that you pump into your car on a cold day may be smaller than the one that you pump on a hot day. Why do you suppose that this is done?
Figure 1.26c shows a double-stranded DNA that is just four pairs of nucleotides long. Human DNA, in contrast, contains a combined total of about 3 billion nucleotide pairs. With each building block being approximately 0.34 nm thick, this gives human DNA an overall length of about 1.0 meter. Its width is about 2.0 nm (\(2 \times 10^{-9}\) m).

\[
3 \times 10^9 \text{ nucleotides} \times \frac{0.34 \text{ nm}}{\text{nucleotide}} \times \frac{1 \times 10^{-9} \text{ m}}{1 \text{ nm}} = 1.0 \text{ m}
\]

To gain an appreciation of the dimensions of human DNA, let us do a quick comparison of its length and width. Dividing the combined length of DNA (1.0 m) by its width (\(2 \times 10^{-9}\) m) shows us that DNA is \(5.0 \times 10^8\) (500 million) times longer than it is wide.

\[
\frac{1.0 \text{ m}}{2.0 \times 10^{-9} \text{ m}} = 5.0 \times 10^8
\]

**Deoxyribonucleic acid (DNA)**

(a) Four different building blocks are used to make DNA. (b) The building blocks are connected to form strands, which (c) combine to form double strands when building blocks interact with one another. (d) The double-stranded DNA twists to form a helix.

(a), (b), and (c) inspired by http://dddxr.blogspot.com/2011/01/aia-maninam-aia-about-human-dna.html; (d) inspired by http://www.biol.unt.edu/~jajohnson/DNA_sequencing_process.
CHAPTER 1  Science and Measurements

What is quite amazing is that this 1 meter stretch of DNA is found in the cell nucleus, which is only 0.024 mm wide. Squeezing this much DNA into a cell nucleus involves a biochemical sleight-of-hand in which DNA gets looped around balls of protein and these wrapped spheres are coiled and coiled again into a very compact shape (Figure 1.27). Compacting DNA in this way allows it to be squeezed into a cell nucleus that is about 80,000 times smaller in diameter than the total length of the DNA.

---

SAMPLE PROBLEM 1.19

The relative length and width of DNA

If a rope that stretched from Los Angeles to New York, a distance of 2462 miles, had the same ratio of length to width as human DNA, how wide (in inches) would it be?

**STRATEGY**
The paragraph directly above showed a calculation for the relative length and width of DNA. Apply this information to the length of 2462 miles.

**SOLUTION**
DNA is \(5.0 \times 10^8\) times longer than it is wide. Dividing 2462 miles by \(5.0 \times 10^8\) gives the width of the rope, in miles:

\[
\frac{2462 \text{ mi}}{5.0 \times 10^8} = 4.9 \times 10^{-6} \text{ mi}
\]

Applying appropriate conversion factors allows us to convert the width into inches.

\[
4.9 \times 10^{-6} \text{ mi} \times \frac{5280 \text{ ft}}{\text{mi}} \times \frac{12 \text{ in}}{\text{ft}} = 0.31 \text{ in}
\]

---

PRACTICE PROBLEM 1.19

Using the same length to width relationship described above, if human DNA were 5.0 inches wide, how long would it be (in kilometers)?

---

**FIGURE 1.27**
Compact DNA  In order to fit into a cell nucleus, which is typically about 80,000 times smaller in diameter than the total length of DNA, the double helix of DNA is wrapped around proteins (blue) and then tightly coiled into chromosomes.
This chapter started with At the Clinic, a brief look at some of the measurements associated with patient treatment. Why does the first chapter of a chemistry textbook begin with this story and then continue with several sections dealing with measurements and units? Because measurements and units are an important part of everyday life. At the grocery store you might wonder which has the better unit price: 32 ounces of apple juice for $1.89 or one-half gallon of the same juice for $3.52. On a road trip, you might pass a sign that reads, “Next gas station 41 miles.” Knowing that your car gets 25 miles to the gallon, is a quarter of a tank of gasoline enough to get you there? A health care professional might be called upon to calculate an IV drip rate (at a drip rate of 150 mL per hour, how many minutes will it take to administer 50 mL of solution?) or drug dosage (if 30 mg is the prescribed dose, how many milliliters should be drawn from a vial containing 50 mg/mL?).

THINKING IT THROUGH
1. What units would typically be used for the different measurements that were made in At the Clinic?
2. List some of the measurements, units, or unit conversion that you might make use of on a typical day.

### CHAPTER 1 OBJECTIVES

<table>
<thead>
<tr>
<th>OBJECTIVE</th>
<th>SUMMARY</th>
<th>SECTION</th>
<th>SAMPLE AND PRACTICE PROBLEMS</th>
<th>END OF CHAPTER PROBLEMS</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.</td>
<td>Explain the terms law, hypothesis, experiment, and theory.</td>
<td>1.1</td>
<td>1.1</td>
<td>1.3–1.6</td>
</tr>
<tr>
<td>2.</td>
<td>Define the terms matter and energy. Describe the three states of matter and the two forms of energy.</td>
<td>1.2</td>
<td>1.3</td>
<td>1.7, 1.8, 1.13–1.26</td>
</tr>
<tr>
<td>3.</td>
<td>Describe and give examples of physical properties and physical change.</td>
<td>1.2</td>
<td>1.2, 1.4</td>
<td>1.9–1.12</td>
</tr>
<tr>
<td>4.</td>
<td>Identify metric, English, and SI units.</td>
<td>1.3</td>
<td>1.5</td>
<td>1.27–1.30</td>
</tr>
<tr>
<td>5.</td>
<td>Express values using scientific notation and metric prefixes.</td>
<td>1.4</td>
<td>1.6, 1.7</td>
<td>1.31–1.44</td>
</tr>
</tbody>
</table>

Scientific laws describe observations but do not attempt to explain them. A theory is a hypothesis (tentative explanation) that has survived repeated testing by experiments.

Matter has mass and occupies space, while energy is the capacity to do work and transfer heat. Matter is typically found as a solid (fixed shape and volume), a liquid (variable shape and fixed volume), or gas (variable shape and volume). Potential energy is stored energy and kinetic energy is the energy of motion.

Potential energy is stored energy and kinetic energy is the energy of motion.

Physical properties, including odor and melting point, and physical changes, including boiling and crushing, can be determined without affecting chemical composition.

Metric units include grams, meters, and liters; English units include pounds, feet, and quarts; and SI units include kilograms, meters, and cubic meters.

In scientific notation, values are expressed as a number between 1 and 10, multiplied by a power of ten \((0.025 = 2.5 \times 10^{-2})\). Metric prefixes are used to create units of different sizes \((2503 \text{ m} = 2.503 \text{ km})\).
OBJECTIVE SUMMARY

6. Describe the difference between the terms accurate and precise.
   Accurate measurements are close to the true value. Precise measurements are grouped closely together.

7. Use the correct number of significant figures to report the results of calculations involving measured quantities.
   Significant figures are all certain digits in a measurement plus the first uncertain one. Exact or defined numbers have no uncertain digits. Results of calculations involving measured quantities must be rounded to the correct number of significant figures (5.5 g + 23.44 g + 0.225 g = 29.2 g and 88 cm × 62.33 cm = 5.5 × 10³ cm²).

8. Identify conversion factors and use them to convert from one unit to another.
   Converting between units may require the use of a conversion factor, which is based on a relationship between units. The conversion factor to use is the one that cancels the original unit(s).

9. Explain the terms density, specific gravity, and specific heat.
   Density is mass of substance in a given volume and specific gravity is the density of a substance divided by the density of water. Specific heat is the amount of heat required to raise the temperature of 1 gram of a substance by 1ºC.

10. Recognize the difference between general chemistry, organic chemistry, and biochemistry.
    General chemistry is a study of the fundamental aspects of chemistry. Organic chemistry is a study of the chemistry of carbon, and biochemistry is a study of the chemistry of living things.

END OF CHAPTER PROBLEMS

Answers to problems whose numbers are printed in color are given in Appendix C. More challenging questions are marked with an asterisk

1.1 Which drawing correctly shows the relationship between pounds and kilograms?

1.2 Which drawing correctly shows the relationship between decigrams and grams?

1.1 The Scientific Method

1.3 Is the statement “What goes up must come down” a scientific law or a scientific theory? Explain.

1.4 Centuries before any experiments had been carried out, philosophers had proposed the existence of the atom. Why are the proposals of these philosophers not considered theories?

1.5 How is a theory different from a hypothesis?

1.6 How is a law different from a hypothesis?

1.2 Matter and Energy

1.7 Define the terms “matter” and “energy.”

1.8 What are the three states of matter?
On a hot day, a glass of iced tea is placed on a table.

- What are some of the physical properties of the ice?
- What change in physical state would you expect to take place if the iced tea sits in the sun for a while?

List a few of the physical properties of a piece of copper wire.

- Give examples of some of the physical changes that a piece of copper wire could undergo.

Give an example of a physical change that involves starting with a liquid and ending up with a gas.

Give an example of a chemical change (change in chemical composition) that involves starting with a liquid and ending up with a gas.

What is potential energy?

- What is kinetic energy?

Describe a situation where an object’s potential energy varies as a result of changes in its position.

Describe a situation where an object’s potential energy varies as a result of changes to its chemical composition.

A battery-powered remote control toy car sits at the bottom of a hill. The car begins to move and is steered up the hill.

- Describe the changes to the car’s kinetic energy.
- Describe the changes to the car’s potential energy that are related to its position.

In the autumn, a leaf falls from a tree.

- Describe changes to the leaf’s potential energy that are related to its position.

Suppose that you are camping in the winter. To obtain drinking water, you use a propane-fueled camp stove to melt snow.

- Is the melting of snow a physical change? Explain.
- When propane burns, the gases carbon dioxide and water vapor are formed. Is the burning of propane a physical change? Explain.
- Describe the potential energy change that takes place for propane as it burns in the stove.
- Describe the kinetic energy change that takes place for water as the snow melts.

Rather than melting snow on a camp stove as described in the previous problem, you decide to eat handfuls of snow.

- Describe the change in the potential energy of your body as the snow that you have swallowed is melted.
- If you are stranded in the woods during the winter, why is it better to obtain water by melting snow than eating it?

What is heat of fusion?

- What is heat of vaporization?

What change in physical state takes place during sublimation?

If you immerse your arm in a bucket of ice water, your arm gets cold. Where does the heat energy from your arm go and what process is the energy used for?

Some over-the-counter (nonprescription) wart removers contain ether. When a few drops are placed on a wart, it feels cold as the ether rapidly evaporates. Where does the heat energy from the wart go and what process is the energy used for?

True or false? If heat is continually added to a pan of boiling water, the temperature of the water continually rises until all of the water has boiled away.

True or false? A container holding a mixture of ice and water at a temperature of 0°C (the freezing temperature of water) is placed in a −30°C freezer. The ice-water mixture stays at a temperature of 0°C until all of the water has frozen.

**1.3 Units of Measurement**

Based on your experience or the information in Table 1.1, which is larger?

- 1 yd or 1 m
- 1 lb or 1 g

Based on your experience or the information in Table 1.1, which is larger?

- 1 pint or 1 L
- 1°F or 1°C

Based on your experience or the information in Table 1.2, which is larger?

- 1 mg or 1 μg
- 1 T or 1 tsp

Based on your experience or the information in Table 1.2, which is larger?

- 1 tsp or 1 fl oz
- 1 gal or 1 L

Convert each number into scientific notation.

- 1.300
- 6,901,000

Convert each number into scientific notation.

- 2,000,000,000
- 850

Convert each number into ordinary notation.

- 7 × 10⁻²
- 7 × 10²

Convert each number into ordinary notation.

- 4.23 × 10⁻⁴
- 4.23 × 10¹
1.46 A scientist buys a thermometer and tests it by measuring the melting point of a particular substance.  
   a. If the melting point of the substance is 1.5°C, what results would show that the thermometer is not precise? 
   b. What results would show that the thermometer is not accurate?

1.47 How many significant figures does each number have? Assume that each is a measured value.  
   a. 1000.0005  
   b. 887.60  
   c. 0.00045  
   d. 0.668  
   e. 70.  

1.48 How many significant figures does each number have? Assume that each is a measured value.  
   a. 1.46  
   b. 3.5985  
   c. 600.2  
   d. 2.0 × 10^3  
   e. 4.55 × 10^3

1.49 Solve each calculation, reporting each answer with the correct number of significant figures. Assume that each value is a measured quantity.  
   a. 14 × 3.6  
   b. 8.0027 ÷ 6.7784  
   c. (1.2 × 10^3 ÷ 0.66) + 1.0  
   d. 12.567 + 34  

1.50 Solve each calculation, reporting each answer with the correct number of significant figures. Assume that each value is a measured quantity.  
   a. 0.114 × 5.2377  
   b. 3.11 × 14.5  
   c. 123.667 – 78.9  
   d. (6.21 + 0.04) × 16.72

1.51 A microbiologist wants to know the circumference of a cell being viewed through a microscope. Estimating the diameter of the cell to be 11 mm and knowing that circumference = \( \pi \times \text{diameter} \) (we will assume that the cell is round, even though that is usually not the case), the microbiologist uses a calculator and gets the answer 34.55751919 \( \mu \text{m} \). Taking significant figures into account, what answer should actually be reported? \( (\pi = 3.141592654 \ldots ) \)

1.52 Given that area = \( \pi \times \text{radius}^2 \) and radius = diameter/2, what is the area of the cell described in Problem 1.51? Report your answer with the correct number of significant figures.

### 1.6 Conversion Factors and the Factor Label Method

1.53 Give the two conversion factors that are based on each equality.  
   a. 12 eggs = 1 dozen  
   b. 1 × 10^3 m = 1 km  
   c. 0.946 L = 1 qt

1.54 Give the two conversion factors that are based on each equality.  
   a. 2 T = 1 oz  
   b. 15 gtt = 1 mL  
   c. 1 mg = 1000 \( \mu \text{g} \)

1.55 Convert  
   a. 48 eggs into dozen  
   b. 250 m into kilometers

1.56 Convert  
   a. 15 T into fluid ounces  
   b. 45 gtt into milliliters
Ivermectin is used to treat dogs that have intestinal parasites. The effective dosage of this drug is 10.5 mg/kg of body weight. How much ivermectin should be given to a 90 kg dog?

Chloroquine is used to treat malaria. Studies have shown that an effective dose for children is 3.5 mg per kilogram (3.5 mg/kg) of body weight, every 6 hours. If a child weighs 12 kg, how many milligrams of this drug should be given in a 24 hour period?

The tranquilizer Valium is sold in 2.0 mL ampoules that contain 50.0 mg of drug per 1.0 mL of liquid (50.0 mg/1.0 mL). If a physician prescribes 25 mg of this drug, how many milliliters should be administered?

An antibiotic is sold in 3.0 mL ampoules that contain 60.0 mg of drug (60.0 mg/3.0 mL). How many milliliters of the antibiotic should be withdrawn from the ampoule if 45 mg are to be administered to a patient?

A vial contains 25 mg/mL of a particular drug. To administer 15 mg of the drug, how many milliliters should be drawn from the vial?

A patient is to receive 50 cc of a drug mixture intravenously over a 1 hr time period. What is the appropriate IV drip rate in gtt/min?

A prescription of antibiotics for a 30 lb child says to give 100 mg three times daily. Is this dosage safe if the proper pediatric dosage range for this drug is 10–30 mg per kilogram of body weight per day?

A patient’s cough syrup prescription comes in a 250 mL bottle. For how long will the cough syrup last if he takes two teaspoons three times a day?

A dose of 3 mg/kg/day (3 mg of drug per kilogram of body weight) of phenobarbital is to be given to a 24 kg patient once a day. Phenobarbital is sold in 35 mg tablets. How many tablets (rounded to the nearest one tablet) should be given to the patient per day?

To treat migraines, valproic acid can be given at a dosage of 15 mg/kg/day (15 mg of drug per kilogram of body weight per day). Valproic acid is sold in 250 mg capsules. How many capsules per day should a 115 lb patient be prescribed?

The antipsychotic drug thioridazine is administered at 0.5 mg/kg/day in three divided doses. The drug is sold in 10 mg tablets. How many tablets should be given per dose to a 180 lb patient?

A prescription calls for giving a 95 lb patient 5 mg/kg/day (5 mg of drug per kilogram of body weight per day) of an anticonvulsant drug, with the half the dose given in the morning and the other half at night. The drug is sold in 100 mg tablets. How many tablets (rounded to the near one tablet) should be given at any given time?
1.81 At 20°C what is the volume in milliliters occupied by 
(see Table 1.7)
a. 15.2 g of water?
b. 2.0 kg of kerosene?
c. \(9.2 \times 10^{-2}\) g of isopropyl alcohol?
d. 75 g of chloroform?
1.82 At 20°C what is the volume in milliliters occupied by 
(see Table 1.7)
a. 1.50 kg of water?
b. 77.2 g of kerosene?
c. 5.0 mg of isopropyl alcohol?
d. 1.0 lb of water?
1.83 A patient has 25.0 mL of blood drawn and this 
volume of blood has a mass of 26.5 g. What is the 
density of the blood?
1.84 A patient has 0.050 L of blood drawn and this volume 
of blood has a mass of 55.0 g. What is the density of 
the blood?
1.85 What is the specific gravity of whole blood at 20°C? 
(See Table 1.7.)
1.86 What is the specific gravity of kerosene at 20°C? (See 
Table 1.7.)
* 1.87 Calculate the number of calories of heat energy 
required for each. (See Table 1.8.)
a. warming 35.0 g of water from 21.0°C to 29.0°C 
b. warming 17.5 g of water from 18.0°C to 54.0°C
* 1.88 Calculate the number of calories of heat energy 
required for each. (See Table 1.8.)
a. warming 2.60 g of isopropyl alcohol from 15.0°C 
to 35.0°C 
b. warming 17.5 g of isopropyl alcohol from 32.0°C 
to 87.0°C
* 1.89 Calculate the number of calories of heat energy 
required for each. (See Table 1.8.)
a. warming 35.0 mL of water from 21.0°C to 29.0°C 
b. warming 17.5 mL of water from 18.0°C to 54.0°C
* 1.90 Calculate the number of calories of heat energy 
required for each. (See Table 1.8.)
a. warming 2.60 mL of isopropyl alcohol from 21.0°F 
to 29.0°F 
b. warming 17.5 mL of isopropyl alcohol from 18.0°F 
to 54.0°F
1.91 How much will the temperature change when 750 g 
of each of the following materials absorbs \(1.25 \times 10^4\) 
cal of heat energy?
a. iron (specific heat = 0.11 cal/g °C)
b. stainless steel (specific heat = 0.12 cal/g °C)
c. aluminum (specific heat = 0.89 J/g °C)
1.92 How much will the temperature change when 55.0 g 
of each of the following materials absorbs 125 cal of heat 
energy?
a. silver (specific heat = 0.056 cal/g °C)
b. olive oil (specific heat = 0.47 cal/g °C)
c. table salt (specific heat = 0.86 J/g °C)
HealthLink | SCIENCE AND MEDICINE

1.101 In the past 200 years, in what ways have scientific discoveries led to changes in the treatment of diabetes?

1.102 Until the late 1980s, what was the source of the insulin used to treat diabetes?

HealthLink | BODY MASS INDEX

1.103 a. A 6’2” tall adult weighs 180 lbs. What is his BMI? Based on this value, what is his status: underweight, normal, overweight, or obese?
   b. Answer part a, but using your height and weight.
   c. A woman stands 1.65 m tall and weighs 72.7 kg. What is her BMI and what is her status?

1.104 September 2006 was the first time that models were banned from a top-level fashion show for being too thin. The organizers of the Madrid Fashion Week defined “too thin” as having a BMI of less than 18. How much would a 5’2” model weigh if she had a BMI of 16?

HealthLink | BODY TEMPERATURE

1.105 A patient has a temperature of 31°C. Should her clinician be concerned?

1.106 Suppose that you take your temperature orally and see that it is 99.1°F. Does this necessarily mean that you are running a fever? Explain.

HealthLink | MAKING WEIGHT

1.107 One of the rule changes that the NCAA made to discourage rapid weight loss was to shorten the time between weigh in and competition from 24 hours to just 2 hours. Why would this discourage athletes from trying to make weight?

1.108 Why does a high urine specific gravity indicate dehydration?

1.109 a. Use the density of water (1.00 g/mL) to derive a conversion factor for water that has the units lb/cup.
   b. If an athlete reduces her body’s water volume by 5.5 cups through restricting fluid intake and sweating in a sauna, how much weight has she lost? Is this a good idea? Explain.

1.110 a. Use the density of water (1.00 g/mL) to derive a conversion factor for water that has the units kg/L.
   b. If an athlete reduces his body’s water volume by 0.75 L through restricting fluid intake and sweating in a sauna, how much weight (in kilograms) has he lost?

Learning Group Problems

1.111 a. Write the two conversion factors that are based on the equality 1 grain = 325 milligrams.
   b. Which conversion factor in your answer to part a would be used to convert grains to milligrams?

   c. One aspirin tablet contains 5.0 grains of aspirin. How many milligrams of aspirin are in two tablets?
   d. How many grams of aspirin are in two tablets?
   e. How many micrograms of aspirin are in two tablets?

1.112 a. Write the two conversion factors that are based on the equality 15 drops = 1 milliliter.
   b. Which conversion factor in your answer to part a would be used to convert milliliters to drops?
   c. 65 drops of water is how many milliliters?
   d. 65 drops of water is how many microliters?
   e. 65 drops of water is how many tablespoons? (See Table 1.2.)

Solutions to Practice Problems

1.1 The battery might be dead or your car might be out of gas. To see if the battery is the problem, you might try turning on the headlights. To see if the car is out of gas, you could either check the gas gauge or add gasoline and try to start the car.

1.2 A chemical change. Even though you begin with a solid and a liquid and end up with a gas, this is not a physical change. Baking soda, vinegar, and carbon dioxide have different chemical compositions.

1.3 a. Based on position, the cup of coffee held at shoulder level; b. Based on temperature (heat is a form of kinetic energy), the cup of hot coffee.

1.4 Water

   a. 273°C; b. 32°C; c. 0°F

   1.6 0.000001 L = 1 × 10⁻⁶ L; 5,000,000 cells = 5 × 10⁶ cells

   1.7 a. 0.02 m = 2 × 10⁻² m = 2 cm; b. 0.0000002 m = 2 × 10⁻⁵ m = 2 mm

   1.8 a. 7.032 × 10³ cal; b. 8.80 × 10¹ J; c. 5 × 10⁻⁵ g; d. 6.430 × 10⁻² lb

   1.9 a. 10,938,000; b. 10,940,000; c. 10,900,000; d. 11,000,000; e. 10,000,000

   1.10 a. 26,100 or 2.61 × 10⁴ (3 significant figures); b. 3.65; c. 300 or 3 × 10² (1 significant figure); d. 8

   1.11 a. 543.0; b. 5.77 × 10⁵; c. 146; d. 46

   1.12 a. 0.12 mg; b. 0.30 mL

   1.13 Yes; the dose was 12.5 mg.

   1.14 4.1 K = −269.1°C = −452.4°F

   1.15 a. 341 g; b. 335 g

   1.16 a. 1.9°C; b. 11°C; c. 1.7°C

   1.17 89.5 cm³ of lead

   1.18 a. 5.9 lb; b. As the answer to part a shows, one gallon of gasoline at 98°C has less mass (contains less matter) than one gallon of gasoline at 8°C. Adjusting the amount of gasoline actually delivered ensures that you will pump the same mass of gasoline, regardless of the temperature.

   1.19 66,000 km