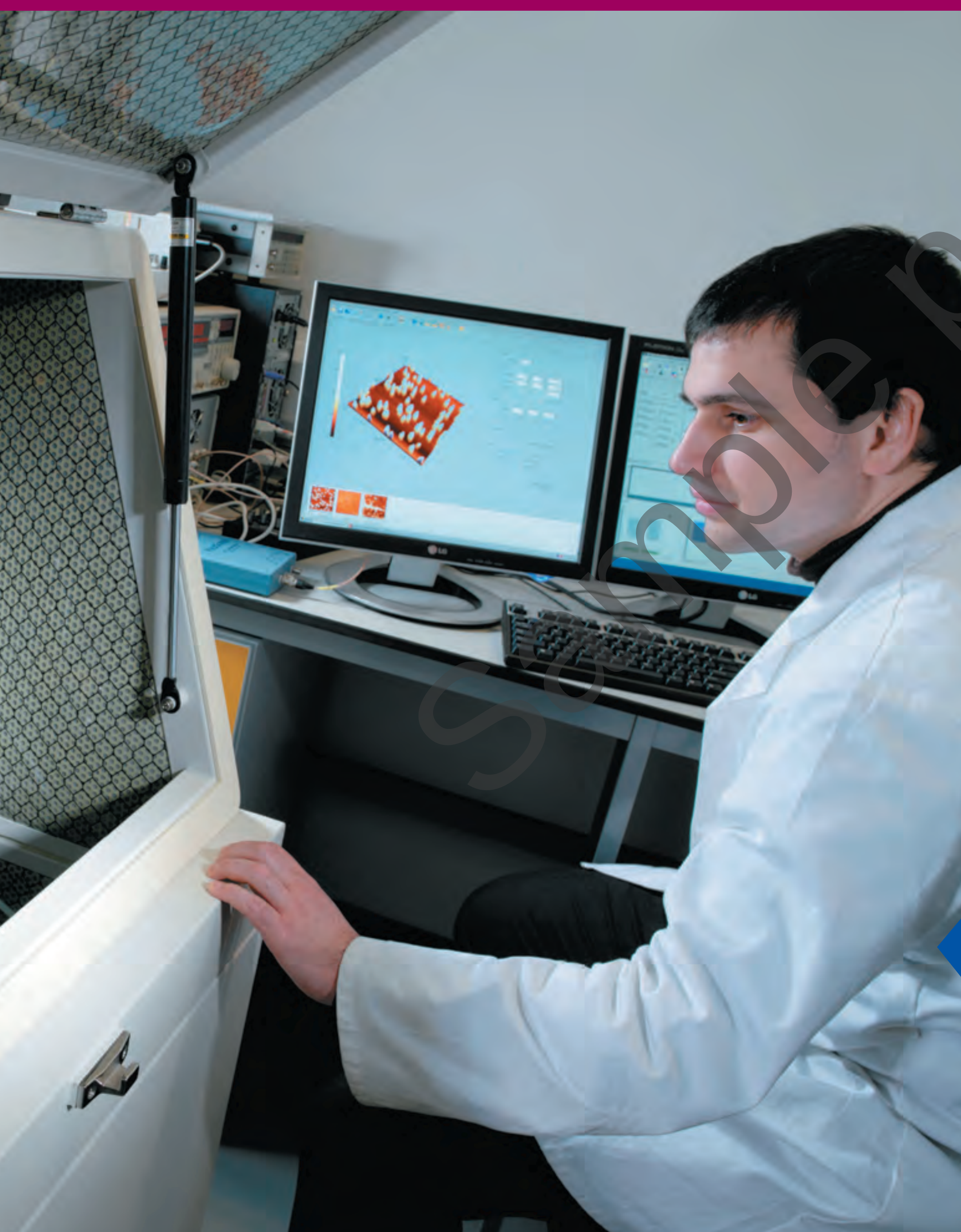
# chapter 1

# Chemical Tools: Experimentation and Measurement



## Contents



Instruments for scientific measurements have changed greatly over the centuries. Modern technology has enabled scientists to make images of extremely tiny particles, even individual atoms, using instruments like this atomic force microscope.

**What are the unique properties of nanoscale (1 nm = $10^{-9}$ m) materials?**

The answer to this question can be found on page 57 in the INQUIRY ?

▲ The sequence of the approximately 5.8 billion nucleic acid units, or *nucleotides*, present in the human genome has been determined using instruments like this automated DNA sequencer.

Life has changed more in the past two centuries than in all the previously recorded span of human history. The Earth's population has increased sevenfold since 1800, and life expectancy has nearly doubled because of our ability to synthesize medicines, control diseases, and increase crop yields. Methods of transportation have changed from horses and buggies to automobiles and airplanes because of our ability to harness the energy in petroleum. Many goods are now made of polymers and ceramics instead of wood and metal because of our ability to manufacture materials with properties unlike any found in nature.

In one way or another, all these changes involve **chemistry**, the study of the composition, properties, and transformations of matter. Chemistry is deeply involved in both the changes that take place in nature and the profound social changes of the past two centuries. In addition, chemistry is central to the current revolution in molecular biology that is revealing the details of how life is genetically regulated. No educated person today can understand the modern world without a basic knowledge of chemistry.

## 1.1 THE SCIENTIFIC METHOD: NANOPARTICLE CATALYSTS FOR FUEL CELLS

By opening this book, you have already decided that you need to know more about chemistry to pursue your future goals. Perhaps you want to learn how living organisms function, how medicines are made, how human activities change the environment, or how alternative fuels produce clean energy. A good place to start is by learning the experimental approach used by scientists to make new discoveries. Do not worry if you do not understand all the details of the chemistry yet, as our focus is on the process of modern interdisciplinary research.

Let's examine a nanoscience application to illustrate the scientific method and how chemical principles are applied to make materials with novel properties. **Nanoscience** is the production and study of structures that have at least one dimension between 1 and 100 nm, where one nanometer is one billionth of a meter. Research on nanomaterials is a fast-growing, multidisciplinary enterprise spanning the fields of chemistry, physics, biology, medicine, materials science, and engineering. Inorganic crystals that have nanoscale dimensions exhibit different properties than bulk material as described in more detail the Inquiry section of this chapter. The properties depend on the size of the particle and can be tuned for applications such as tools for diagnosing and treating disease or platforms for sustainable energy.

**LOOKING AHEAD . . .**
The rates of chemical reactions and how they are increased by **catalysts** are described in Chapter 14.

**LOOKING AHEAD . . .**
Chapter 4 describes different types of reactions including **redox reactions** that involve a transfer of electrons. We'll see in Chapter 19 how redox reactions can be used to generate electricity in a **fuel cell.**

One research area is the use of nanoparticle catalysts for reactions occurring in fuel cells. A **catalyst** is a substance that speeds up the rate of a chemical reaction. A **fuel cell** is a device that uses a fuel such as hydrogen to produce electricity. Fuel cells operate much like a battery, but they require a continuous input of fuel. Two reactions occur at two different electrodes in a hydrogen fuel cell. At one electrode, hydrogen ($H_2$) is converted to protons ($H^+$), and at the other electrode, oxygen ($O_2$) reacts with protons to produce water. The reactions in the fuel cell involve a transfer of electrons and are called **redox reactions.** The electrons produced by reaction 1 (below) travel through a wire and are used in reaction 2. The movement of electrons through a wire generates electricity. A fuel cell is considered to be *zero emission* because the overall reaction of hydrogen with oxygen produces electricity but pure water is the only product.

$$\begin{array}{ll} \text{Reaction 1:} & 2\,H_2(g) \longrightarrow 4\,H^+(aq) + 4\,e^- \\ \text{Reaction 2:} & O_2(g) + 4\,H^+(aq) + 4\,e^- \longrightarrow 2\,H_2O(l) \\ \hline \text{Overall reaction:} & 2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(l) \end{array}$$

Fuel cells are a promising technology in the quest for a carbon-neutral energy economy, but one obstacle to their use is the slow rate of conversion of oxygen to water in reaction 2. Platinum particles coated on the surface of the electrode have been used as a catalyst to speed up the reaction, but platinum is very expensive. Nanoparticles made from palladium alloys have shown promise as a cost-effective alternative catalyst. An **alloy** is a mixture of metals, and therefore a palladium alloy is a mixture of palladium (Pd) and some other metal such as copper (Cu).

In order to develop a useful catalyst for hydrogen fuel cells, chemists apply the scientific method to carefully control different characteristics of PdCu nanoparticles and measure their effect on the rate of the oxygen reaction. Some characteristics of nanoparticles that can be varied are relative amounts of palladium and copper, the size of the particles, and the shape of the particles. Amazingly nanoparticles exist in a variety of shapes including spheres, cubes, and octopods (**FIGURE 1.1**)!

**BIG IDEA Question 1**

What is an obstacle to the widespread use of hydrogen fuel cells, and how can nanoparticles be used to overcome the problem?

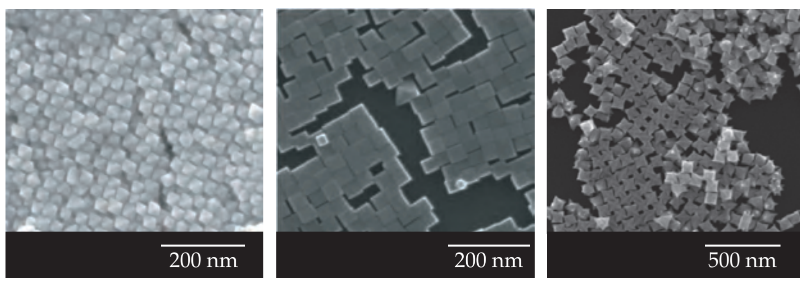



**FIGURE 1.1**

**From left to right, scanning electron microscopy images of octahedral gold nanoparticles, cubic palladium nanoparticles, and eight-branched gold-palladium nanoparticles called "octopods."** Note that the orientation of the octopodal nanoparticles allows only four of the branches to be viewed at a time.

Images courtesy of the Skrabalak research group at Indiana University.

## The Scientific Method

A general approach to research is called the **scientific method.** The scientific method is an iterative process involving the formulation of questions arising from observations, careful design of experiments, and thoughtful analysis of results. The scientific method involves identifying ways to test the validity of new ideas, and seldom is there only one way to go about it. The main elements of the scientific method, outlined in **FIGURE 1.2,** are the following:

- ***Observation.*** **Observations** are a systematic recording of natural phenomena and may be **qualitative,** descriptive in nature, or **quantitative,** involving measurements.
- ***Hypothesis.*** A **hypothesis** is a possible explanation for the observation developed based upon facts collected from previous experiments as well as scientific knowledge



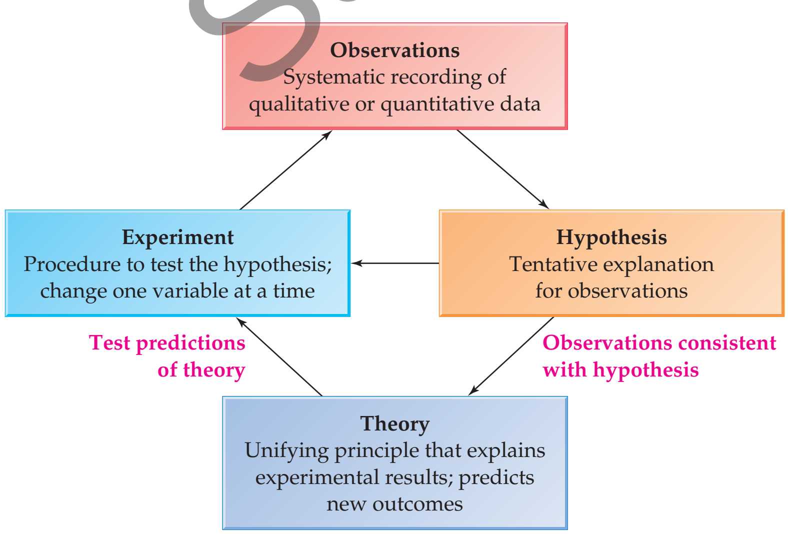



**FIGURE 1.2**

**The scientific method.** An iteractive experimental approach is used in scientific research. Hypotheses and theories are refined based on new experiments and obervations.

**Figure It out**

What is developed when numerous experimental observations support a hypothesis?

**Answer:** Theory

and intuition. The hypothesis may not be correct, but it must be testable with an experiment.

- ***Experiment.*** An **experiment** is a procedure for testing the hypothesis. Experiments are most useful when they are performed in a *controlled* manner, meaning that only one variable is changed at a time while all others remain constant.
- ***Theory.*** A **theory** is developed from a hypothesis consistent with experimental data and is a unifying principle that explains experimental results. It also makes predictions about related systems, and new experiments are carried out to verify the theory.

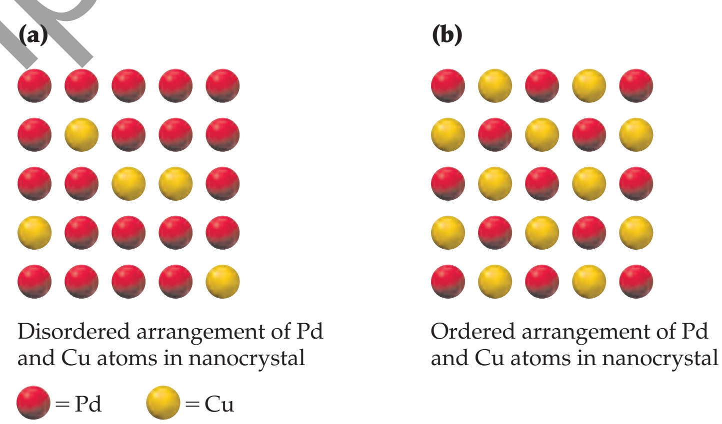
Keep in mind as you study chemistry or any other science that theories can never be absolutely proven. There's always the chance that a new experiment might give results that can't be explained by present theory. All a theory can do is provide the best explanation that we can come up with at the present time. Science is an ever-changing field where new observations are made with increasingly sophisticated equipment; it is always possible that existing theories may be modified in the future.

Scientific research begins with a driving question that is frequently based on experimental observations or a desire to learn about the unknown. In the case of PdCu nanoparticles, an observed increase in the fuel cell reaction rate led to the question "How can variables related to size, shape, and composition of nanoparticles be controlled to optimize catalytic activity?" Professor Sara Skrabalak at Indiana University leads a team of scientists researching methods for synthesizing high-quality nanomaterials, where these variables are precisely controlled. Although previous research projects involving numerous techniques attempted to control the size, shape, and composition of the nanoparticles, the distributions of palladium and copper atoms in the crystals were found to be statistically random. **FIGURE 1.3a** illustrates a random or disordered arrangement of Pd and Cu atoms in a crystal. **FIGURE 1.3b** illustrates an ordered arrangement with a pattern of alternating Pd and Cu atoms. Without fixed arrangements of atoms, it is impossible to correlate chemical structure with properties such as catalytic activity.



▶ **FIGURE 1.3**

**Simple schematic of the arrangement of Pd and Cu atoms in a nanocrystal.** (a) A disordered arrangement of Pd and Cu atoms does not have a repeating pattern. (b) An ordered arrangement of atoms has the repeating pattern of alternating Pd and Cu atoms.




The general hypothesis that the Skrabalak group tested was that larger PdCu nanoparticles with lower surface energies would facilitate the transition from disordered to ordered structures. Student researchers carefully controlled the rate of particle growth by depositing palladium and copper on the surface of a smaller particle. Then various imaging techniques were used to elucidate the atomic-level structure of the nanoparticles and measure their size distribution. Electron microscopy data revealed a thin shell of Pd over an ordered PdCu core called the B2 phase. **FIGURE 1.4a** shows a

(a) (b)

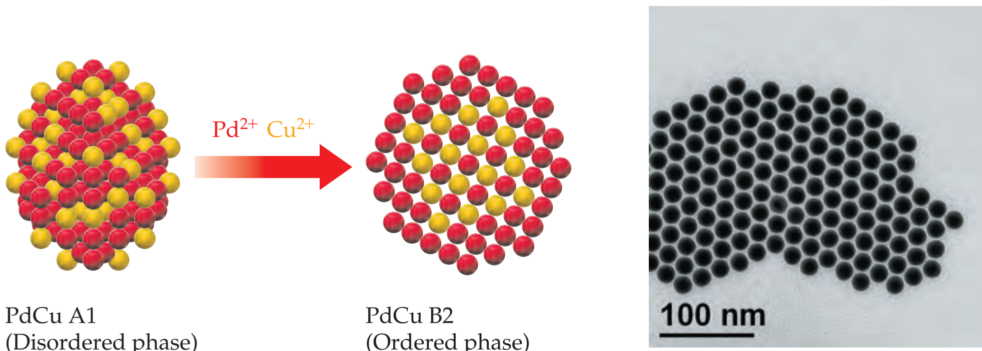



Image courtesy of the Skrabalak research group at Indiana University.

**FIGURE 1.4**
(**a**) Reaction scheme showing the disordered A1 phase in the PdCu nanoparticle being converted to the ordered B2 phase. (**b**) Transmission electron microscopy image of the B2 phase in the PdCu nanoparticles showing spherical particles of uniform size.

**Figure It Out**

Use the scale bar in Figure 1.4b to determine the approximate mean diameter of the PdCu nanoparticles.

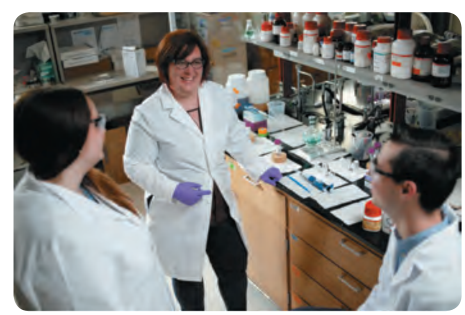
**Answer:** 20 nm

random distribution of Pd and Cu atoms in the A1 phase that was converted by new synthesis methods to the ordered B2 phase. **FIGURE 1.4b** shows a transmission electron microscope image of PdCu nanoparticles in the B2 phase. The spherical particles have a uniform size distribution with a mean diameter of 18.9 nm.

Many iterations of the scientific method were used by the researchers to devise controlled synthesis techniques for PdCu nanoparticles. Observations from experiments led to new hypotheses and additional experiments to test them. Once studies of the growth mechanism enabled reproducible synthesis of ordered PdCu nanoparticles, they were tested for catalytic activity. The ordered nanoparticles exhibited superior catalytic activity in increasing the rate of oxygen reaction in the fuel cell when compared with PdCu nanoparticles with disordered structures. In summary, Professor Skrabalak's research on nanomaterial synthesis leads to the design of better nanoparticle catalysts for fuel cells and other applications.

Many different chemical principles that you will learn about in this book are central to the design of nanomaterials. In Chapter 8, Bonding Theories and Molecular Structure, you will learn about bonds and forces that cause atoms to aggregate into nanoparticles. Chapter 4, Reactions in Aqueous Solutions, describes how to calculate solution concentrations important in synthetic techniques. Rates of reactions and factors that influence them are explored in Chapter 14, Kinetics. In Chapter 19, on electrochemistry, redox reactions central to forming nanoparticles are described.

At universities around the world, students participate in research projects like the one on nanoparticle synthesis and characterization just described. It is the authors' sincere hope that by reading this book you can gain an appreciation for how chemistry is used in solving many of the world's problems and you become competent with the essential chemical principles needed to contribute to important research projects.



▲ Professor Sara Skrabalak in the lab with undergraduate student researchers working on the synthesis of nanoparticles.
Photo courtesy of Indiana University.



## 1.2 MEASUREMENTS: SI UNITS AND SCIENTIFIC NOTATION

Chemistry is an experimental science. But if our experiments are to be reproducible, we must be able to fully describe the substances we're working with—their amounts, volumes, temperatures, and so forth. Thus, one of the most important requirements in chemistry is that we have a way to measure things.

Under an international agreement concluded in 1960, scientists throughout the world now use the International System of Units for measurement, abbreviated **SI unit** for the French *Système Internationale d'Unités*. Based on the metric system, which is used in all industrialized countries of the world except the United States, the SI system has seven fundamental units (**TABLE 1.1**). These seven fundamental units, along with others derived from them, suffice for all scientific measurements. We'll look at three of

**BIG IDEA Question 2**

What are the fundamental SI units of measure for mass, length, and temperature?

**TABLE 1.1 The Seven Fundamental SI Units of Measure**

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

the most common units in this chapter—those for mass, length, and temperature—and will discuss others as the need arises in later chapters.

One problem with any system of measurement is that the sizes of the units often turn out to be inconveniently large or small. For example, a chemist describing the diameter of a sodium atom (0.000 000 000 372 m) would find the meter (m) to be inconveniently large, but an astronomer describing the average distance from the Earth to the Sun (150,000,000,000 m) would find the meter to be inconveniently small. For this reason, SI units are modified through the use of prefixes when they refer to either smaller or larger quantities. Thus, the prefix *milli-* means one-thousandth, and a *milli*meter (mm) is 1/1000 of 1 meter. Similarly, the prefix *kilo-* means one thousand, and a *kilo*meter (km) is 1000 meters. (Note that the SI unit for mass [kilogram] already contains the *kilo-* prefix.) A list of prefixes is shown in **TABLE 1.2,** with the most commonly used ones in red.

Notice how numbers that are either very large or very small are indicated in Table 1.2 using an exponential format called **scientific notation.** For example, the number 55,000 is written in scientific notation as $5.5 \times 10^4$ and the number 0.003 20 as $3.20 \times 10^{-3}$.



**TABLE 1.2 Some Prefixes for Multiples of SI Units. Common prefixes and symbols in the chemical sciences are shown in red**

| Factor | Prefix | Symbol | Example |
|---|---|---|---|
| $1{,}000{,}000{,}000{,}000 = 10^{12}$ | tera | T | 1 teragram (Tg) $= 10^{12}$ g |
| $1{,}000{,}000{,}000 = 10^{9}$ | giga | G | 1 gigameter (Gm) $= 10^{9}$ m |
| $1{,}000{,}000 = 10^{6}$ | mega | M | 1 megameter (Mm) $= 10^{6}$ m |
| $1000 = 10^{3}$ | kilo | k | 1 kilogram (kg) $= 10^{3}$ g |
| $100 = 10^{2}$ | hecto | h | 1 hectogram (hg) = 100 g |
| $10 = 10^{1}$ | deka | da | 1 dekagram (dag) = 10 g |
| $0.1 = 10^{-1}$ | deci | d | 1 decimeter (dm) = 0.1 m |
| $0.01 = 10^{-2}$ | centi | c | 1 centimeter (cm) = 0.01 m |
| $0.001 = 10^{-3}$ | milli | m | 1 milligram (mg) = 0.001 g |
| *$0.000\ 001 = 10^{-6}$ | micro | $\mu$ | 1 micrometer ($\mu$m) $= 10^{-6}$ m |
| *$0.000\ 000\ 001 = 10^{-9}$ | nano | n | 1 nanosecond (ns) $= 10^{-9}$ s |
| *$0.000\ 000\ 000\ 001 = 10^{-12}$ | pico | p | 1 picosecond (ps) $= 10^{-12}$ s |
| *$0.000\ 000\ 000\ 000\ 001 = 10^{-15}$ | femto | f | 1 femtomole (fmol) $= 10^{-15}$ mol |

*For very small numbers, it is becoming common in scientific work to leave a thin space every three digits to the right of the decimal point, analogous to the comma placed every three digits to the left of the decimal point in large numbers.

Review Appendix A if you are uncomfortable with scientific notation or if you need to brush up on how to do mathematical manipulations on numbers with exponents.

Notice also that all measurements contain both a number and a unit label. A number alone is not much good without a unit to define it. If you asked a friend how far it was to the nearest tennis court, the answer "3" alone wouldn't tell you much: 3 blocks? 3 kilometers? 3 miles? Worked Example 1.1 explains how to write a number in scientific notation and represent the unit in prefix notation.

## WORKED EXAMPLE 1.1

### Expressing Measurements Using Scientific Notation and SI Units

Express the following quantities in scientific notation and then express the number and unit with the most appropriate prefix.

**(a)** The diameter of a sodium atom, 0.000 000 000 372 m
**(b)** The distance from the Earth to the Sun, 150,000,000,000 m

#### STRATEGY

To write a number in scientific notation, shift the decimal point to the right or left by $n$ places until you obtain a number between 1 and 10. If the decimal is shifted to the right, $n$ is negative, and if the decimal is shifted to the left, $n$ is positive. Then multiply the result by $10^n$. Choose a prefix for the unit that is close to the exponent of the number written in scientific notation.

#### SOLUTION

**(a)**

Number of decimal places point was shifted to the right.

$$0.000\ 000\ 000\ 372\ \text{m} = 3.72 \times 10^{-10}\ \text{m} = 372\ \text{pm}$$

**(b)**

Number of decimal places point was shifted to the left.

$$150{,}000{,}000{,}000\ \text{m} = 1.5 \times 10^{11}\ \text{m} = 150\ \text{Gm}$$

▶ **PRACTICE 1.1** Express the diameter of a nanoparticle (0.000 000 050 m) in scientific notation, and then express the number and unit with the most appropriate prefix.

▶ **APPLY 1.2** Express the following quantities in scientific notation using fundamental SI units of mass and length given in Table 1.1.

**(a)** The diameter of a human hair, 70 $\mu$m
**(b)** The mass of carbon dioxide emitted from a large power plant each year, 20 Tg

All **PRACTICE** and **APPLY** problems are interactive in the eText.



## 1.3 MASS AND ITS MEASUREMENT

**Mass** is defined as the amount of *matter* in an object. **Matter,** in turn, is a catch-all term used to describe anything with a physical presence—anything you can touch, taste, or smell. (Stated more scientifically, matter is anything that has mass.) Mass is measured in SI units by the **kilogram** (**kg**; 1 kg = 2.205 U.S. lb). Because the kilogram is too large for many purposes in chemistry, the metric **gram** (**g**; 1 g = 0.001 kg), the **milligram** (**mg**; 1 mg = 0.001 g = $10^{-6}$ kg), and the **microgram** (**$\mu$g**; 1 $\mu$g = 0.001 mg = $10^{-6}$ g = $10^{-9}$ kg) are more commonly used. (The symbol **$\mu$** is the lowercase Greek letter mu.) One gram is a bit less than half the mass of a new U.S. dime.

▲ The mass of a U.S. dime is approximately 2.27 g.

$$1\ \text{kg} = 1000\ \text{g} = 1{,}000{,}000\ \text{mg} = 1{,}000{,}000{,}000\ \mu\text{g} \qquad (2.205\ \text{lb})$$
$$1\ \text{g} = 1000\ \text{mg} = 1{,}000{,}000\ \mu\text{g} \qquad (0.035\ 27\ \text{oz})$$
$$1\ \text{mg} = 1000\ \mu\text{g}$$

The standard kilogram is set as the mass of a cylindrical bar of platinum–iridium alloy stored in a vault in a suburb of Paris, France. There are 40 copies of this bar distributed throughout the world, with two (Numbers 4 and 20) stored at the U.S. National Institute of Standards and Technology near Washington, D.C.

The terms *mass* and *weight*, although often used interchangeably, have quite different meanings. *Mass* is a physical property that measures the amount of matter in an object, whereas *weight* measures the force with which gravity pulls on an object. Mass is independent of an object's location: your body has the same amount of matter whether you're on Earth or on the moon. Weight, however, *does* depend on an object's location. If you weigh 140 lb on Earth, you would weigh only about 23 lb on the moon, which has a lower gravity than the Earth.

At the same location on Earth, two objects with identical masses experience an identical pull of the Earth's gravity and have identical weights. Thus, the mass of an object can be measured by comparing its weight to the weight of a reference standard of known mass. Much of the confusion between mass and weight is simply due to a language problem. We speak of "weighing" when we really mean that we are measuring mass by comparing two weights. **FIGURE 1.5** shows balances typically used for measuring mass in the laboratory.

**BIG IDEA Question 3**

Which prefix for the unit of grams is most appropriate for reporting the mass of a grain of sand?

▶ **FIGURE 1.5**
**Some balances used for measuring mass in the laboratory.**



## 1.4 LENGTH AND ITS MEASUREMENT

**BIG IDEA Question 4**

Which prefix for the unit of meter is most appropriate for reporting the diameter of a molecule?

The **meter (m)** is the standard unit of length in the SI system. Although originally defined in 1790 as being 1 ten-millionth of the distance from the equator to the North Pole, the meter was redefined in 1889 as the distance between two thin lines on a bar of platinum–iridium alloy stored near Paris, France. To accommodate an increasing need for precision, the meter was redefined again in 1983 as equal to the distance traveled by light through a vacuum in 1/299,792,458 second. Although this new definition isn't as easy to grasp as the distance between two scratches on a bar, it has the great advantage that it can't be lost or damaged.

One meter is 39.37 inches, about 10% longer than an English yard and much too large for most measurements in chemistry. Other more commonly used measures of length are the **centimeter** (**cm**; 1 cm = 0.01 m, a bit less than half an inch), the **millimeter** (**mm**; 1 mm = 0.001 m, about the thickness of a U.S. dime), the **micrometer** (**$\mu$m**; 1 $\mu$m = $10^{-6}$ m), the **nanometer** (**nm**; 1 nm = $10^{-9}$ m), and the **picometer** (**pm**; 1 pm = $10^{-12}$ m). Thus, a chemist might refer to the diameter of a sodium atom as 372 pm ($3.72 \times 10^{-10}$ m).

$$1\ \text{m} = 100\ \text{cm} = 1000\ \text{mm} = 1{,}000{,}000\ \mu\text{m} = 1{,}000{,}000{,}000\ \text{nm} \qquad (1.0936\ \text{yd})$$
$$1\ \text{cm} = 10\ \text{mm} = 10{,}000\ \mu\text{m} = 10{,}000{,}000\ \text{nm} \qquad (0.3937\ \text{in.})$$
$$1\ \text{mm} = 1000\ \mu\text{m} = 1{,}000{,}000\ \text{nm}$$

▲ The length of the bacteria on the tip of this pin is about $5 \times 10^{-7}$ m or 500 nm.

## 1.5 TEMPERATURE AND ITS MEASUREMENT

Just as the kilogram and the meter are slowly replacing the pound and the yard as common units for mass and length measurement in the United States, the **Celsius degree** (°C) is slowly replacing the degree **Fahrenheit** (°F) as the common unit for temperature measurement. In scientific work, however, the **kelvin** (**K**) has replaced both. (Note that we say only "kelvin," not "degree kelvin.")

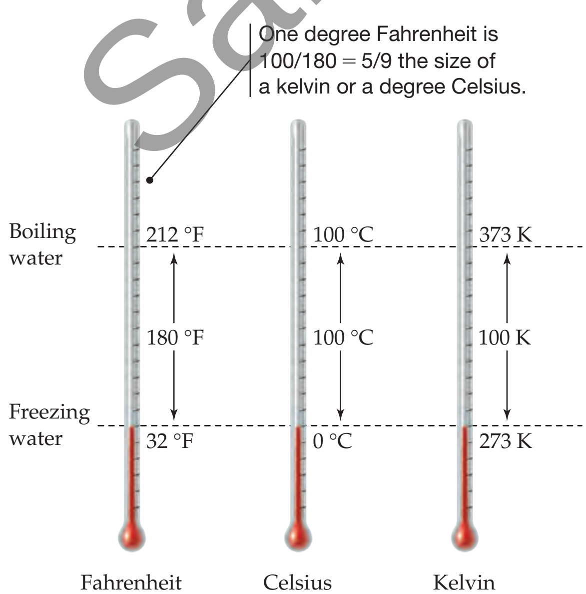
For all practical purposes, the kelvin and the degree Celsius are the same—both are one-hundredth of the interval between the freezing point of water and the boiling point of water at standard atmospheric pressure. The only real difference between the two units is that the numbers assigned to various points on the scales differ. Whereas the Celsius scale assigns a value of 0 °C to the freezing point of water and 100 °C to the boiling point of water, the Kelvin scale assigns a value of 0 K to the coldest possible temperature, −273.15 °C, sometimes called *absolute zero*. Thus, 0 K = −273.15 °C and 273.15 K = 0 °C. For example, a warm spring day with a Celsius temperature of 25 °C has a Kelvin temperature of 25 + 273.15 = 298 K.

**Relationship between the Kelvin and Celsius scales**

$$\text{Temperature in K} = \text{Temperature in °C} + 273.15$$
$$\text{Temperature in °C} = \text{Temperature in K} - 273.15$$

In contrast to the Kelvin and Celsius scales, the common Fahrenheit scale specifies an interval of 180° between the freezing point (32 °F) and the boiling point (212 °F) of water. Thus, it takes 180 degrees Fahrenheit to cover the same range as 100 degrees Celsius (or kelvins), and a degree Fahrenheit is therefore only 100/180 = 5/9 as large as a degree Celsius. **FIGURE 1.6** compares the Fahrenheit, Celsius, and Kelvin scales.

Two adjustments are needed to convert between Fahrenheit and Celsius scales—one to adjust for the difference in degree size and one to adjust for the difference in zero points. The size adjustment is made using the relationships 1 °C = (9/5) °F




◀ **FIGURE 1.6**
**A comparison of the Fahrenheit, Celsius, and Kelvin temperature scales.**

**◀ Figure It Out**

Which represents the largest increase in temperature: +10 °F, +10 °C, or +10 K?

**Answer:** Temperature changes of +10 °C or +10 K are equal and larger than +10 °F.

and 1 °F = (5/9) °C. The zero-point adjustment is made by remembering that the freezing point of water is higher by 32 on the Fahrenheit scale than on the Celsius scale. Thus, if you want to convert from Celsius to Fahrenheit, you do a size adjustment (multiply °C by 9/5) and then a zero-point adjustment (add 32). If you want to convert from Fahrenheit to Celsius, you find out how many Fahrenheit degrees there are above freezing (by subtracting 32) and then do a size adjustment (multiply by 5/9). The following formulas describe the conversions:

**Celsius to Fahrenheit**

$$°\text{F} = \left(\frac{9\ °\text{F}}{5\ \cancel{°\text{C}}} \times \cancel{°\text{C}}\right) + 32\ °\text{F}$$

**Fahrenheit to Celsius**

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

Worked Example 1.2 shows how to convert between temperature scales and estimate the answer. Before tackling Worked Example 1.2, we'd like to point out that the Worked Examples in this book suggest a series of steps useful in organizing and analyzing information.

## Problem-Solving Steps In Worked Examples

**IDENTIFY**

Classify pertinent information as known or unknown. (The quantity needed in the answer will, of course, be unknown.) Specify units and symbols to help identify necessary equations and procedures.

**STRATEGY**

Find a relationship between the known information and unknown answer, and plan a strategy for getting from one to the other.

**SOLUTION**

Solve the problem.

**CHECK**

If possible, make a rough estimate to be sure your calculated answer is reasonable, and think about the number and sign to make sure it makes sense.



### WORKED EXAMPLE 1.2

### Converting between Temperature Scales

The normal body temperature of a healthy adult is 98.6 °F. What is this value on both Celsius and Kelvin scales?

**IDENTIFY**

| Known | Unknown |
|---|---|
| Temperature, 98.6 °F | Temperature in units of °C and K |

**STRATEGY**

Use the formulas for converting Fahrenheit to Celsius and Celsius to Kelvin.

**SOLUTION**

Set up an equation using the temperature conversion formula for changing from Fahrenheit to Celsius:

$$°\text{C} = \left(\frac{5\ °\text{C}}{9\ \cancel{°\text{F}}}\right)(98.6\ \cancel{°\text{F}} - 32\ \cancel{°\text{F}}) = 37.0\ °\text{C}$$

Converting to kelvin gives a temperature of 37.0 + 273.15 = 310.2 K.

**CHECK**

A useful way to double-check a calculation is to estimate the answer. Body temperature in °F is first rounded to the nearest whole number, 99. To account for the difference in zero points of the two scales, 32 is subtracted: 99 − 32 = 67. Because a degree Fahrenheit is only 5/9 as large as a degree Celsius, the next step is to multiply by 5/9, which can be approximated by dividing by two: 67/2 = 33.5. The estimate is only slightly lower than the calculated answer (37.0), indicating the mathematical operations have most likely been performed correctly. Estimating Fahrenheit to Celsius conversions is useful as daily temperatures are reported on these two different scales throughout the world.

**PRACTICE 1.3** The melting point of table salt is 1474 °F. What temperature is this on the Celsius and Kelvin scales?

**APPLY 1.4** The metal gallium has a relatively low melting point for a metal, 302.91 K. If the temperature in the cargo compartment carrying a shipment of gallium has a temperature of 88 °F, is the gallium in the solid or liquid state?

▲ The melting point of sodium chloride is 1474 °F.

## 1.6 DERIVED UNITS: VOLUME AND ITS MEASUREMENT

Look back at the seven fundamental SI units given in Table 1.1, and you'll find that measures for such familiar quantities as area, volume, density, speed, and pressure are missing. All are examples of *derived* quantities rather than fundamental quantities because they can be expressed using one or more of the seven base units (**TABLE 1.3**).

Volume, the amount of space occupied by an object, is measured in SI units by the **cubic meter ($m^3$)**, defined as the amount of space occupied by a cube 1 meter on edge (**FIGURE 1.7**).

A cubic meter equals 264.2 U.S. gallons, much too large a quantity for normal use in chemistry. As a result, smaller, more convenient measures are commonly employed. Both the **cubic decimeter ($dm^3$)** ($1\ dm^3 = 0.001\ m^3$), equal in size to the more familiar metric **liter (L)**, and the **cubic centimeter ($cm^3$)** ($1\ cm^3 = 0.001\ dm^3 = 10^{-6}\ m^3$), equal in size to the metric **milliliter (mL)**, are particularly convenient. Slightly larger than 1 U.S. quart, a liter has the volume of a cube 1 dm on edge. Similarly, a milliliter has the volume of a cube 1 cm on edge (Figure 1.7).

$$1\ m^3 = 1000\ dm^3 = 1{,}000{,}000\ cm^3 \qquad (264.2\ \text{gal})$$

$$1\ dm^3 = 1\text{L} = 1000\ \text{mL} \qquad (1.057\ \text{qt})$$

**FIGURE 1.8** shows some of the equipment frequently used in the laboratory for measuring liquid volume.

**BIG IDEA Question 5**

What is the edge length of a cube with a volume of 1 L?

**TABLE 1.3 Some Derived Quantities**

| Quantity | Definition | Derived Unit (Name) |
|---|---|---|
| Area | Length times length | $m^2$ |
| Volume | Area times length | $m^3$ |
| Density | Mass per unit volume | $kg/m^3$ |
| Speed | Distance per unit time | m/s |
| Acceleration | Change in speed per unit time | $m/s^2$ |
| Force | Mass times acceleration | $(kg{\cdot}m)/s^2$ (newton, N) |
| Pressure | Force per unit area | $kg/(m{\cdot}s^2)$ (pascal, Pa) |
| Energy | Force times distance | $(kg{\cdot}m^2)/s^2$ (joule, J) |

**FIGURE 1.7**

**Units for measuring volume.** A cubic meter is the volume of a cube 1 meter along each edge.

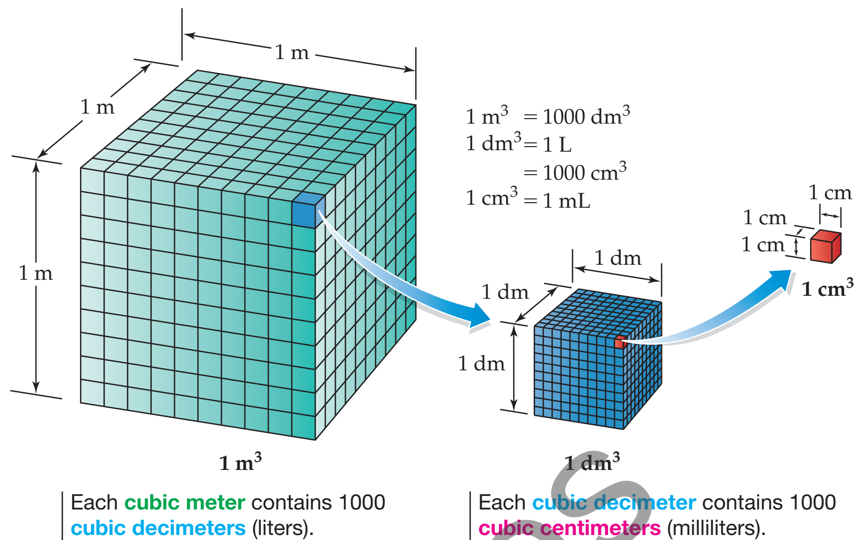





**FIGURE 1.8**

**Common items of laboratory equipment used for measuring liquid volume.**

## 1.7 DERIVED UNITS: DENSITY AND ITS MEASUREMENT

The relationship between the mass of an object and its volume is called *density*. **Density** is calculated as the mass of an object divided by its volume and is expressed in the SI derived unit g/mL for a liquid or g/cm$^3$ for a solid. The densities of some common materials are given in **TABLE 1.4.**

$$\textbf{Density} = \frac{\text{Mass (g)}}{\text{Volume (mL or cm}^3)}$$

**TABLE 1.4 Densities of Some Common Materials**

| Substance | Density (g /cm$^3$) |
|---|---|
| Ice (0 °C) | 0.917 |
| Water (3.98 °C) | 1.0000 |
| Gold | 19.31 |
| Helium (25 °C) | 0.000 164 |
| Air (25 °C) | 0.001 185 |
| Human fat | 0.94 |
| Human muscle | 1.06 |
| Cork | 0.22–0.26 |
| Balsa wood | 0.12 |
| Earth | 5.54 |

Because most substances change in volume when heated or cooled, densities are temperature dependent. At 3.98 °C for example, a 1.0000 mL container holds exactly 1.0000 g of water (density = 1.0000 g/mL). As the temperature is raised, however, the volume occupied by the water expands so that only 0.9584 g fits in the 1.0000 mL container at 100 °C (density = 0.9584 g/mL). *When reporting a density, the temperature must also be specified.*

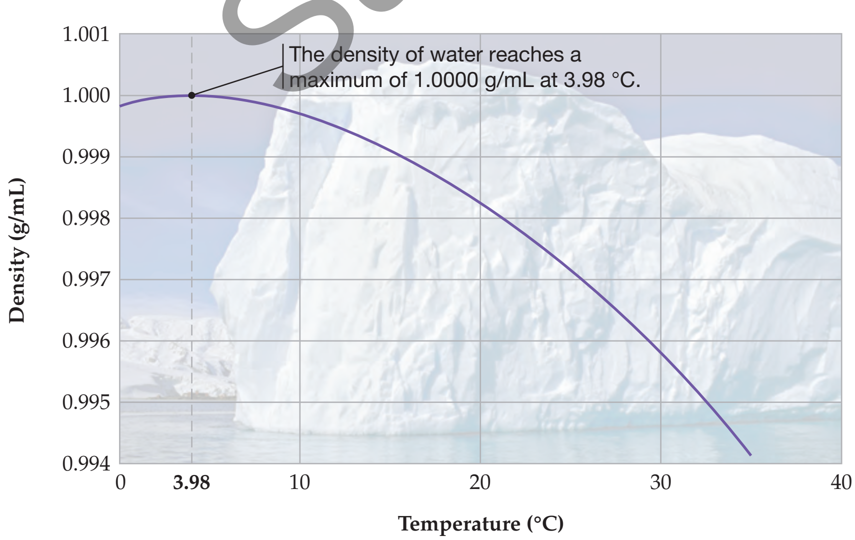
Although most substances expand when heated and contract when cooled, water behaves differently. Water contracts when cooled from 100 °C to 3.98 °C, but below this temperature it begins to expand again. Thus, the density of liquid water is at its maximum of 1.0000 g/mL at 3.98 °C but decreases to 0.999 87 g/mL at 0 °C (**FIGURE 1.9**). When freezing occurs, the density drops still further to a value of 0.917 g/cm$^3$ for ice at 0 °C. Ice and any other substance with a density less than that of water will float, but any substance with a density greater than that of water will sink.

Knowing the density of a substance, particularly a liquid, can be very useful because it's often easier to measure a liquid by volume than by mass. Suppose, for example, that you needed 1.55 g of ethyl alcohol. Rather than trying to weigh exactly the right amount, it would be much easier to look up the density of ethyl alcohol (0.7893 g/mL at 20 °C) and measure the correct volume with a syringe as shown in Figure 1.8.

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} \quad \text{so} \quad \text{Volume} = \frac{\text{Mass}}{\text{Density}}$$

$$\text{Volume} = \frac{1.55\ \cancel{\text{g}}\ \text{ethyl alcohol}}{0.7893\ \frac{\cancel{\text{g}}}{\text{mL}}} = 1.96\ \text{mL ethyl alcohol}$$

▲ Which weighs more, the brass weight or the pillow? Actually, both have identical masses and weights, but the brass has a higher density because its volume is smaller.




◀ **FIGURE 1.9**
**The density of water at different temperatures.**

**◀ Figure It Out**

Would 10.0 mL of water at 10 °C have the same mass as 10.0 mL of water at 25 °C?

**Answer:** No, the density of water changes with temperature. 10.0 mL of water at 10 °C would have a higher mass than 10.0 mL of water at 25 °C because the density is higher at 10 °C.

### WORKED EXAMPLE 1.3

#### Relating Mass and Volume Using Density

What is the volume in $cm^3$ of 201 g of gold?

**IDENTIFY**

| Known | Unknown |
|---|---|
| Mass of gold (201 g) | Volume of gold ($cm^3$) |

**STRATEGY**

Rearrange the equation for density to solve for volume. Use the known value for the density of gold in Table 1.4 and the mass of gold in the equation.

**SOLUTION**

$$\text{Volume} = \frac{201\ \cancel{\text{g}}\ \text{gold}}{19.31\ \cancel{\text{g}}\ /\text{cm}^3} = 10.4\ \text{cm}^3\ \text{gold}$$

**CHECK**

The mass of gold (201 g) is roughly 10 times larger than the density (19.31 g), so the estimate for volume is 10 $cm^3$, which agrees with the calculated answer.

**PRACTICE 1.5** Chloroform, a substance once used as an anesthetic, has a density of 1.483 g/mL at 20 °C. How many milliliters would you use if you needed 9.37 g? (1 mL = 1 $cm^3$)

**APPLY 1.6** You are beachcombing on summer vacation and find a silver bracelet. You take it to the jeweler, and he tells you that it is silver plated and will give you $10 for it. You do not want to be swindled, so you take the bracelet to your chemistry lab and find its mass on a balance (80.0 g). To measure the volume, you place the bracelet in a graduated cylinder (Figure 1.8) containing 10.0 mL of water at 20 °C. The final volume in the graduated cylinder after the bracelet has been added is 17.61 mL. The density of silver at 20 °C is 10.5 g/$cm^3$ and 1 $cm^3$ = 1 mL. What can you conclude about the identity of the metal in the bracelet?

▲ How might you determine if this bracelet is pure silver?



## 1.8 DERIVED UNITS: ENERGY AND ITS MEASUREMENT

The word *energy* is familiar to everyone but is surprisingly hard to define in simple, nontechnical terms. A good working definition, however, is to say that **energy** is the capacity to supply heat or do work. The water falling over a dam, for instance, contains energy that can be used to turn a turbine and generate electricity. A tank of propane gas contains energy that, when released in the chemical process of combustion, can heat a house or barbecue a hamburger.

Energy is classified as either *kinetic* or *potential*. **Kinetic energy** ($E_K$) is the energy of motion. The amount of kinetic energy in a moving object with mass $m$ and velocity $v$ is given by the equation

$$\boxed{E_K = \frac{1}{2}mv^2}$$

The larger the mass of an object and the larger its velocity, the larger the amount of kinetic energy. Thus, water that has fallen over a dam from a great height has a greater velocity and more kinetic energy than the same amount of water that has fallen only a short distance.

**BIG IDEA Question 6**

Which of the following statements describes potential energy?

(a) Atoms in a crystal vibrate at temperatures above 0 K.

(b) A negatively charged electron is attracted to a positively charged nucleus.

(c) At room temperature, the average speed of an oxygen molecule is approximately 500 m/s.

**Potential energy** ($E_p$), by contrast, is stored energy—perhaps stored in an object because of its height or in a molecule because of chemical reactions it can undergo. The water sitting in a reservoir behind the dam contains potential energy because of its height above the stream at the bottom of the dam. When the water is allowed to fall, its potential energy is converted into kinetic energy. Propane and other substances used as fuels contain potential energy because they can undergo a combustion reaction with oxygen that releases **energy** as heat and work.

The units for energy, $(\text{kg·m}^2)/\text{s}^2$, follow from the expression for kinetic energy, $E_K = 1/2mv^2$. If, for instance, your body has a mass of 50.0 kg (about 110 lb) and you are riding a bicycle at a velocity of 10.0 m/s (about 22 mi/h), your kinetic energy is 2500 $(\text{kg·m}^2)/\text{s}^2$.

$$E_K = \frac{1}{2}mv^2 = \frac{1}{2}(50.0\ \text{kg})\left(10.0\,\frac{\text{m}}{\text{s}}\right)^2 = 2500\,\frac{\text{kg·m}^2}{\text{s}^2} = 2500\ \text{J}$$

The SI derived unit for energy $(\text{kg·m}^2)/\text{s}^2$ is given the name **joule** (J) after the English physicist James Prescott Joule (1818–1889). The joule is a fairly small amount of energy—it takes roughly 100,000 J to heat a coffee cup full of water from room temperature to boiling—so kilojoules (kJ) are more frequently used in chemistry.

In addition to the SI energy unit joule, some chemists and biochemists still use the unit calorie (cal, with a lowercase c). Originally defined as the amount of energy necessary to raise the temperature of 1 g of water by 1 °C (specifically, from 14.5 °C to 15.5 °C), one calorie is now defined as exactly 4.184 J.

$$1\ \text{cal} = 4.184\ \text{J (exactly)}$$

Nutritionists use the somewhat confusing unit Calorie (Cal, with a capital *C*), which is equal to 1000 calories, or 1 kilocalorie (kcal).

$$1\ \text{Cal} = 1000\ \text{cal} = 1\ \text{kcal} = 4.184\ \text{kJ}$$

The energy value, or caloric content, of food is measured in Calories. Thus, the statement that a banana contains 70 Calories means that 70 Cal (70 kcal, or 290 kJ) of energy is released when the banana is used by the body for fuel.

**LOOKING AHEAD . . .**
In Chapter 9, we'll calculate the amount of **energy** released or absorbed during a chemical reaction.

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

▲ (**a**) A 75-watt incandescent bulb uses energy at the rate of 75 J/s. Only about 5% of that energy appears as light, however; the remaining 95% is given off as heat. (**b**) Energy-efficient light-emitting diode (LED) bulbs are replacing incandescent lights.



## WORKED EXAMPLE 1.4

### Calculating Kinetic Energy

(**a**) What is the kinetic energy in joules of a 2360 lb (1070 kg) car moving at 63.3 mi/h (28.3 m/s)? Express the number in scientific notation.
(**b**) Express the number and unit using an appropriate prefix.

**IDENTIFY**

| Known | Unknown |
| --- | --- |
| Mass (1070 kg), velocity (28.3 m/s) | Kinetic energy ($E_K$) in units of joules (J) |

**STRATEGY**

Use the formula to calculate kinetic energy. If mass is in units of kg and velocity is in units of m/s, then energy will be calculated in units of joules because 1 joule $= \dfrac{1\ \text{kg·m}^2}{\text{s}^2}$.

**SOLUTION**

(**a**) $$E_K = \frac{1}{2}mv^2 = \frac{1}{2}(1070\ \text{kg})\left(28.3\,\frac{\text{m}}{\text{s}}\right)^2 = 428{,}476\,\frac{\text{kg·m}^2}{\text{s}^2}$$
$$= 4.28476 \times 10^5\ \text{J}$$

(**b**) Kilo ($10^3$) and mega ($10^6$) are both prefixes with exponents similar to the answer in part (**a**). Therefore, both 428.476 kJ and 0.428 476 MJ are reasonable answers.

**CHECK**

An answer with a large magnitude should be expected as a car has a very large mass. To estimate the magnitude of kinetic energy, express mass and velocity in the equation using exponents only.

$$E_K \approx (10^3\,\text{kg})\left(\frac{10^1\,\text{m}}{\text{s}}\right)^2 = 10^5\,\frac{\text{kg·m}^2}{\text{s}^2}$$

The estimated power of 10 agrees with the detailed calculation.

*continued on next page*

**PRACTICE 1.7** Some radioactive materials emit a type of radiation called alpha particles at high velocity.

What is the kinetic energy in joules of an alpha particle with a mass of $6.6 \times 10^{-24}$ g and a speed of $1.5 \times 10^{7}$ m/s? Express the number in scientific notation.

**APPLY 1.8** A baseball with a mass of 450 g has a kinetic energy of 406 J. Calculate the velocity of the baseball in units of m/s.

## 1.9 ACCURACY, PRECISION, AND SIGNIFICANT FIGURES IN MEASUREMENT

Measuring things, whether in cooking, construction, or chemistry, is something that most of us do every day. But how good are those measurements? Any measurement is only as good as the skill of the person doing the work and the reliability of the equipment being used. You've probably noticed, for instance, that you often get slightly different readings when you weigh yourself on a bathroom scale and on a scale at the doctor's office, so there's always some uncertainty about your real weight. The same is true in chemistry—there is always some uncertainty in the value of a measurement.

In talking about the degree of uncertainty in a measurement, we use the words *accuracy* and *precision*. Although most of us use the words interchangeably in daily life, there's actually an important distinction between them. **Accuracy** refers to how close to the true value a given measurement is, whereas **precision** refers to how well a number of independent measurements agree with one another. To see the difference, imagine that you weigh a tennis ball whose true mass is 54.441 778 g. Assume that you take three independent measurements on each of three different types of balance to obtain the data shown in the following table.


▲ This tennis ball has a mass of about 54 g.

| Measurement # | Bathroom Scale | Lab Balance | Analytical Balance |
|---|---|---|---|
| 1 | 0.1 kg | 54.4 g | 54.4418 g |
| 2 | 0.0 kg | 54.5 g | 54.4417 g |
| 3 | 0.1 kg | 54.3 g | 54.4418 g |
| (average) | (0.07 kg) | (54.4 g) | (54.4418 g) |

If you use a bathroom scale, your measurement (average = 0.07 kg) is neither accurate nor precise. Its accuracy is poor because it measures to only one digit that is far from the true value, and its precision is poor because any two measurements may differ substantially. If you now weigh the ball on an inexpensive laboratory balance, the value you get (average = 54.4 g) has three digits and is fairly accurate, but it is still not very precise because the three readings vary from 54.3 g to 54.5 g, perhaps due to air movements in the room or a sticky mechanism. Finally, if you weigh the ball on an expensive analytical balance like those found in research laboratories, your measurement (average = 54.4418 g) is both precise and accurate. It's accurate because the measurement is very close to the true value, and it's precise because it has six digits that vary little from one reading to another.

To indicate the uncertainty in a measurement, *the value you record should use all the digits you are sure of plus one additional digit that you estimate*. In reading a thermometer that has a mark for each degree, for example, you could be certain about the digits of the nearest mark—say, 25 °C—but you would have to estimate between two marks—say, between 25 °C and 26 °C—to obtain a value of 25.3 °C.

The total number of digits recorded for a measurement is called the measurement's number of **significant figures.** For example, the mass of the tennis ball as determined on the single-pan balance (54.4 g) has three significant figures, whereas the mass