Exploring Density

Materials
three beakers, a 250 mL graduated cylinder, a colorless and clear cylindrical container, shampoo, water, rubbing alcohol, red food coloring, a medicine dropper, a spoon or stirring rod, and small solid objects (such as a paper clip, plastic soda-bottle cap, aluminum nail, or piece of cork).

Procedure
1. Pour about 250 mL each of shampoo, water, and rubbing alcohol into separate beakers.
2. Stir a drop of food coloring into the rubbing alcohol and the shampoo.
3. Carefully pour each liquid (in the following order: shampoo, water, rubbing alcohol) down the inside of the colorless cylindrical container to a depth of 2.5 cm (about 1 in.).
4. Record the positions of the three layers of liquids.
5. Gently add each small solid object to the layered liquids and record its resting location.

Think About It
1. Propose an explanation for the order of layering of the liquids.
2. Do you think it would make any difference if you added the liquids in a different order? Why or why not?
3.1 Measurements and Their Uncertainty

Connecting to Your World

On January 4, 2004, the Mars Exploration Rover Spirit landed on Mars. Equipped with five scientific instruments and a rock abrasion tool (shown at left), Spirit was sent to examine the Martian surface around Gusev Crater, a wide basin that may have once held a lake. Each day of its mission, Spirit recorded measurements for analysis. This data helped scientists learn about the geology and climate on Mars. All measurements have some uncertainty. In the chemistry laboratory, you must strive for accuracy and precision in your measurements.

Using and Expressing Measurements

Your height (67 inches), your weight (134 pounds), and the speed you drive on the highway (65 miles/hour) are some familiar examples of measurements. A measurement is a quantity that has both a number and a unit. Everyone makes and uses measurements. For instance, you decide how to dress in the morning based on the temperature outside. If you were baking cookies, you would measure the volumes of the ingredients as indicated in the recipe.

Such everyday situations are similar to those faced by scientists. Measurements are fundamental to the experimental sciences. For that reason, it is important to be able to make measurements and to decide whether a measurement is correct. The units typically used in the sciences are those of the International System of Measurements (SI).

In chemistry, you will often encounter very large or very small numbers. A single gram of hydrogen, for example, contains approximately 602,000,000,000,000,000,000,000,000 hydrogen atoms. The mass of an atom of gold is 0.000 000 000 000 000 000 000 327 gram. Writing and using such large and small numbers is very cumbersome. You can work more easily with these numbers by writing them in scientific, or exponential, notation.

In scientific notation, a given number is written as the product of two numbers: a coefficient and 10 raised to a power. For example, the number 602,000,000,000,000,000,000,000,000 written in scientific notation is \(6.02 \times 10^{23}\). The coefficient in this number is 6.02. In scientific notation, the coefficient is always a number equal to or greater than one and less than ten. The power of 10, or exponent, in this example is 23. Figure 3.1 illustrates how to express the number of stars in a galaxy by using scientific notation. For more practice on writing numbers in scientific notation, refer to page R56 of Appendix C.

Guide for Reading

Key Concepts

- How do measurements relate to science?
- How do you evaluate accuracy and precision?
- Why must measurements be reported to the correct number of significant figures?
- How does the precision of a calculated answer compare to the precision of the measurements used to obtain it?

Vocabulary
measurement
scientific notation
accuracy
precision
accepted value
experimental value
error
percent error
significant figures

Reading Strategy

Building Vocabulary  As you read, write a definition of each key term in your own words.

\[200,000,000,000 = 2 \times 10^{11}\]
Decimal moves 11 places to the left.

Figure 3.1 Expressing very large numbers, such as the estimated number of stars in a galaxy, is easier if scientific notation is used.
Figure 3.2 The distribution of darts illustrates the difference between accuracy and precision. 

- **a** Good accuracy and good precision: The darts are close to the bull’s-eye and to one another.  
- **b** Poor accuracy and good precision: The darts are far from the bull’s-eye but close to one another.  
- **c** Poor accuracy and poor precision: The darts are far from the bull’s-eye and from one another.

**Accuracy, Precision, and Error**

Your success in the chemistry lab and in many of your daily activities depends on your ability to make reliable measurements. Ideally, measurements should be both correct and reproducible.

**Accuracy and Precision** Correctness and reproducibility relate to the concepts of accuracy and precision, two words that mean the same thing to many people. In chemistry, however, their meanings are quite different. **Accuracy** is a measure of how close a measurement comes to the actual or true value of whatever is measured. **Precision** is a measure of how close a series of measurements are to one another. **To evaluate the accuracy of a measurement, the measured value must be compared to the correct value. To evaluate the precision of a measurement, you must compare the values of two or more repeated measurements.**

Darts on a dartboard illustrate accuracy and precision in measurement. Let the bull’s-eye of the dartboard represent the true, or correct, value of what you are measuring. The closeness of a dart to the bull’s-eye corresponds to the degree of accuracy. The closer it comes to the bull’s-eye, the more accurately the dart was thrown. The closeness of several darts to one another corresponds to the degree of precision. The closer together the darts are, the greater the precision and the reproducibility.

Look at Figure 3.2 as you consider the following outcomes:

- **a.** All of the darts land close to the bull’s-eye and to one another. Closeness to the bull’s-eye means that the degree of accuracy is great. Each dart in the bull’s-eye corresponds to an accurate measurement of a value. Closeness of the darts to one another indicates high precision.

- **b.** All of the darts land close to one another but far from the bull’s-eye. The precision is high because of the closeness of grouping and thus the high level of reproducibility. The results are inaccurate, however, because of the distance of the darts from the bull’s-eye.

- **c.** The darts land far from one another and from the bull’s-eye. The results are both inaccurate and imprecise.

**Checkpoint** How does accuracy differ from precision?
Determining Error  Note that an individual measurement may be accurate or inaccurate. Suppose you use a thermometer to measure the boiling point of pure water at standard pressure. The thermometer reads 99.1°C. You probably know that the true or accepted value of the boiling point of pure water under these conditions is actually 100.0°C. There is a difference between the accepted value, which is the correct value based on reliable references, and the experimental value, the value measured in the lab. The difference between the experimental value and the accepted value is called the error.

\[
\text{Error} = \text{experimental value} - \text{accepted value}
\]

Error can be positive or negative depending on whether the experimental value is greater than or less than the accepted value.

For the boiling-point measurement, the error is 99.1°C − 100.0°C, or −0.9°C. The magnitude of the error shows the amount by which the experimental value differs from the accepted value. Often, it is useful to calculate the relative error, or percent error. The percent error is the absolute value of the error divided by the accepted value, multiplied by 100%.

\[
\text{Percent error} = \left| \frac{\text{error}}{\text{accepted value}} \right| \times 100\%
\]

Using the absolute value of the error means that the percent error will always be a positive value. For the boiling-point measurement, the percent error is calculated as follows.

\[
\text{Percent error} = \left| \frac{99.1°C - 100.0°C}{100.0°C} \right| \times 100\%
\]

\[
= \frac{0.9°C}{100.0°C} \times 100\%
\]

\[
= 0.009 \times 100\%
\]

\[
= 0.9\%
\]

Just because a measuring device works doesn’t necessarily mean that it is accurate. As Figure 3.3 shows, a weighing scale that does not read zero when nothing is on it is bound to yield error. In order to weigh yourself accurately, you must first make sure that the scale is zeroed.

Figure 3.3 The scale below has not been properly zeroed, so the reading obtained for the person’s weight is inaccurate. There is a difference between the person’s correct weight and the measured value. Calculating What is the percent error of a measured value of 114 lb if the person’s actual weight is 107 lb?
Significant Figures in Measurements

Supermarkets often provide scales like the one in Figure 3.4. Customers use these scales to measure the weight of produce that is priced per pound. If you use a scale that is calibrated in 0.1-lb intervals, you can easily read the scale to the nearest tenth of a pound. With such a scale, however, you can also estimate the weight to the nearest hundredth of a pound by noting the position of the pointer between calibration marks.

Suppose you estimate a weight that lies between 2.4 lb and 2.5 lb to be 2.46 lb. The number in this estimated measurement has three digits. The first two digits in the measurement (2 and 4) are known with certainty. But the rightmost digit (6) has been estimated and involves some uncertainty. These three reported digits all convey useful information, however, and are called significant figures. The significant figures in a measurement include all of the digits that are known, plus a last digit that is estimated. Measurements must always be reported to the correct number of significant figures because calculated answers often depend on the number of significant figures in the values used in the calculation.

Instruments differ in the number of significant figures that can be obtained from their use and thus in the precision of measurements. The three meter sticks in Figure 3.5 can be used to make successively more precise measurements of the board.
3.5 Zeros at the rightmost end of a measurement that lie to the left of an understood decimal point are not significant if they serve as placeholders to show the magnitude of the number. The zeros in the measurements 300 meters, 7000 meters, and 27,210 meters are not significant. The numbers of significant figures in these values are one, one, and four, respectively. If such zeros were known measured values, however, then they would be significant. For example, if all of the zeros in the measurement 300 meters were significant, writing the value in scientific notation as \(3.00 \times 10^2\) meters makes it clear that these zeros are significant.

There are two situations in which numbers have an unlimited number of significant figures. The first involves counting. If you count 23 people in your classroom, then there are exactly 23 people, and this value has an unlimited number of significant figures. The second situation involves exactly defined quantities such as those found within a system of measurement. When, for example, you write 60 min = 1 hr, or 100 cm = 1 m, each of these numbers has an unlimited number of significant figures. As you shall soon see, exact quantities do not affect the process of rounding an answer to the correct number of significant figures.
CONCEPTUAL PROBLEM 3.1

Counting Significant Figures in Measurements
How many significant figures are in each measurement?

a. 123 m  
b. 40,506 mm
c. 9.8000 × 10^4 m  
d. 22 meter sticks
e. 0.070 80 m  
f. 98,000 m

1 Analyze Identify the relevant concepts.
The location of each zero in the measurement and the location of the decimal point determine which of the rules apply for determining significant figures.

2 Solve Apply the concepts to this problem.
All nonzero digits are significant (rule 1). Use rules 2 through 6 to determine if the zeros are significant.

a. three (rule 1)  
b. five (rule 2)
c. five (rule 4)  
d. unlimited (rule 6)
e. four (rules 2, 3, 4)  
f. two (rule 5)

Practice Problems

1. Count the significant figures in each length.
   a. 0.057 30 meters  
b. 8765 meters
c. 0.000 73 meters  
d. 40.007 meters

2. How many significant figures are in each measurement?
   a. 143 grams  
b. 0.074 meter
c. 8.750 × 10^-2 gram  
d. 1.072 meter

Interactive Textbook

Problem-Solving 3.2 Solve Problem 2 with the help of an interactive guided tutorial.

with ChemASAP

Significant Figures in Calculations
Suppose you use a calculator to find the area of a floor that measures 7.7 meters by 5.4 meters. The calculator would give an answer of 41.58 square meters. The calculated area is expressed to four significant figures. However, each of the measurements used in the calculation is expressed to only two significant figures. So the answer must also be reported to two significant figures (42 m²). In general, a calculated answer cannot be more precise than the least precise measurement from which it was calculated. The calculated value must be rounded to make it consistent with the measurements from which it was calculated.

Rounding To round a number, you must first decide how many significant figures the answer should have. This decision depends on the given measurements and on the mathematical process used to arrive at the answer. Once you know the number of significant figures your answer should have, round to that many digits, counting from the left. If the digit immediately to the right of the last significant digit is less than 5, it is simply dropped and the value of the last significant digit stays the same. If the digit in question is 5 or greater, the value of the digit in the last significant place is increased by 1.

Checkpoint Why must a calculated answer generally be rounded?
SAMPLE PROBLEM 3.1

Rounding Measurements

Round off each measurement to the number of significant figures shown in parentheses. Write the answers in scientific notation.

a. 314.721 meters (four)
b. 0.001 775 meter (two)
c. 8792 meters (two)

1 Analyze Identify the relevant concepts.
Round off each measurement to the number of significant figures indicated. Then apply the rules for expressing numbers in scientific notation.

2 Solve Apply the concepts to this problem.
Count from the left and apply the rule to the digit immediately to the right of the digit to which you are rounding. The arrow points to the digit immediately following the last significant digit.

a. 314.721 meters
   2 is less than 5, so you do not round up.
   314.7 meters = $3.147 \times 10^2$ meters

b. 0.001 775 meter
   7 is greater than 5, so round up.
   0.0018 meter = $1.8 \times 10^{-3}$ meter

c. 8792 meters
   9 is greater than 5, so round up
   8800 meters = $8.8 \times 10^3$ meters

3 Evaluate Do the results make sense?
The rules for rounding and for writing numbers in scientific notation have been correctly applied.

Practice Problems

3. Round each measurement to three significant figures. Write your answers in scientific notation.
   a. 87.073 meters
   b. $4.3621 \times 10^8$ meters
   c. 0.01552 meter
   d. 9009 meters
   e. $1.7777 \times 10^{-2}$ meter
   f. 629.55 meters

4. Round each measurement in Practice Problem 3 to one significant figure. Write each of your answers in scientific notation.
**Addition and Subtraction** The answer to an addition or subtraction calculation should be rounded to the same number of decimal places (not digits) as the measurement with the least number of decimal places. Work through Sample Problem 3.2 below which provides an example of rounding in an addition calculation.

### SAMPLE PROBLEM 3.2

**Significant Figures in Addition**

Calculate the sum of the three measurements. Give the answer to the correct number of significant figures.

$$12.52 \text{ meters} + 349.0 \text{ meters} + 8.24 \text{ meters}$$

1. **Analyze** Identify the relevant concepts.
   Calculate the sum and then analyze each measurement to determine the number of decimal places required in the answer.

2. **Solve** Apply the concepts to this problem.
   Align the decimal points and add the numbers. Round the answer to match the measurement with the least number of decimal places.

   $\begin{align*}
   &12.52 \text{ meters} \\
   + &349.0 \text{ meters} \\
   + &8.24 \text{ meters} \\
   \hline
   &369.76 \text{ meters}
   \end{align*}$

   The second measurement (349.0 meters) has the least number of digits (one) to the right of the decimal point. Thus the answer must be rounded to one digit after the decimal point. The answer is rounded to 369.8 meters, or $3.698 \times 10^2$ meters.

3. **Evaluate** Does the result make sense?
   The mathematical operation has been correctly carried out and the resulting answer is reported to the correct number of decimal places.

### Practice Problems

5. Perform each operation. Express your answers to the correct number of significant figures.
   - a. $61.2 \text{ meters} + 9.35 \text{ meters} + 8.6 \text{ meters}$
   - b. $9.44 \text{ meters} - 2.11 \text{ meters}$
   - c. $1.36 \text{ meters} + 10.17 \text{ meters}$
   - d. $34.61 \text{ meters} - 17.3 \text{ meters}$

6. Find the total mass of three diamonds that have masses of 14.2 grams, 8.73 grams, and 0.912 gram.
**Multiplication and Division** In calculations involving multiplication and division, you need to round the answer to the same number of significant figures as the measurement with the least number of significant figures. The position of the decimal point has nothing to do with the rounding process when multiplying and dividing measurements. The position of the decimal point is important only in rounding the answers of addition or subtraction problems.

**Checkpoint** 
How many significant figures must you round an answer to when performing multiplication or division?

**SAMPLE PROBLEM 3.3**  
**Significant Figures in Multiplication and Division**  
Perform the following operations. Give the answers to the correct number of significant figures.  
a. 7.55 meters \( \times \) 0.34 meter  
b. 2.10 meters \( \times \) 0.70 meter  
c. 2.4526 meters \( \div \) 8.4

1. **Analyze**  Identify the relevant concepts.  
Perform the required math operation and then analyze each of the original numbers to determine the correct number of significant figures required in the answer.

2. **Solve**  Apply the concepts to this problem.  
Round the answers to match the measurement with the least number of significant figures.  
a. 7.55 meters \( \times \) 0.34 meter = 2.567 (meter)\(^2\) = 2.6 meters\(^2\)  
   (0.34 meter has two significant figures)  
b. 2.10 meters \( \times \) 0.70 meter = 1.47 (meter)\(^2\) = 1.5 meters\(^2\)  
   (0.70 meter has two significant figures)  
c. 2.4526 meters \( \div \) 8.4 = 0.291 976 meter = 0.29 meter  
   (8.4 has two significant figures)

3. **Evaluate**  Do the results make sense?  
The mathematical operations have been performed correctly, and the resulting answers are reported to the correct number of places.

**Practice Problems**
7. Solve each problem. Give your answers to the correct number of significant figures and in scientific notation.  
a. 8.3 meters \( \times \) 2.22 meters  
b. 8432 meters \( \div \) 12.5  
c. 35.2 seconds \( \times \) \( \frac{1 \text{ minute}}{60 \text{ seconds}} \)
8. Calculate the volume of a warehouse that has inside dimensions of 22.4 meters by 11.3 meters by 5.2 meters.  
(Volume = \( l \times w \times h \))
Quick LAB

Accuracy and Precision

Purpose
To measure the dimensions of an object as accurately and precisely as possible and to apply rules for rounding answers calculated from the measurements.

Materials
- 3 inch × 5 inch index card
- metric ruler

Procedure
1. Use a metric ruler to measure in centimeters the length and width of an index card as accurately and precisely as you can. The hundredths place in your measurement should be estimated.

2. Calculate the perimeter \(2 \times (\text{length} + \text{width})\) and the area \(\text{length} \times \text{width}\) of the index card. Write both your unrounded answers and your correctly rounded answers on the chalkboard.

3. How many significant figures are in your calculated value for the area? In your calculated value for the perimeter? Do your rounded answers have as many significant figures as your classmates' measurements?

4. Assume that the correct (accurate) length and width of the card are 12.70 cm and 7.62 cm, respectively. Calculate the percent error for each of your two measurements.

3.1 Section Assessment

9. 🟢 Key Concept How do measurements relate to experimental science?

10. 🟢 Key Concept How are accuracy and precision evaluated?

11. 🟢 Key Concept Why must a given measurement always be reported to the correct number of significant figures?

12. 🟢 Key Concept How does the precision of a calculated answer compare to the precision of the measurements used to obtain it?

13. A technician experimentally determined the boiling point of octane to be 124.1°C. The actual boiling point of octane is 125.7°C. Calculate the error and the percent error.

14. Determine the number of significant figures in each of the following.
   a. 11 soccer players
   b. 0.070 020 meter
   c. 10,800 meters
   d. 5.00 cubic meters

15. Solve the following and express each answer in scientific notation and to the correct number of significant figures.
   a. \((5.3 \times 10^6) + (1.3 \times 10^6)\)
   b. \((7.2 \times 10^{-4}) ÷ (1.8 \times 10^3)\)
   c. \(10^5 \times 10^{-2} \times 10^6\)
   d. \((9.12 \times 10^{-3}) - (4.7 \times 10^{-2})\)
   e. \((5.4 \times 10^6) \times (3.5 \times 10^6)\)

Explanatory Paragraph
Explain the differences between the accuracy, precision, and error of a measurement.

Interactive Textbook
Assessment 3.1 Test yourself on the concepts in Section 3.1.

with ChemASAP
3.2 The International System of Units

Connecting to Your World

"Are we there yet?" You may have asked this question during a long road trip with family or friends. To find out how much farther you have to go, you can read the roadside signs that list destinations and their distances. In the signs shown here, however, the distances are listed as numbers with no units attached. Is Carrieton 44 kilometers or 44 miles away? Without the units, you can't be sure. When you make a measurement, you must assign the correct units to the numerical value. Without the units, it is impossible to communicate the measurement clearly to others.

Measuring with SI Units

All measurements depend on units that serve as reference standards. The standards of measurement used in science are those of the metric system. The metric system is important because of its simplicity and ease of use. All metric units are based on multiples of 10. As a result, you can convert between units easily. The metric system was originally established in France in 1795. The International System of Units (abbreviated SI, after the French name, *Le Système International d’Unités*) is a revised version of the metric system. The SI was adopted by international agreement in 1960. There are seven SI base units, which are listed in Table 3.1. From these base units, all other SI units of measurement can be derived. The five SI base units commonly used by chemists are the meter, the kilogram, the kelvin, the second, and the mole.

All measured quantities can be reported in SI units. Sometimes, however, non-SI units are preferred for convenience or for practical reasons. In this textbook you will learn about both SI and non-SI units.

| Table 3.1 |
|------------------|-----------------|-------|
| **Quantity**     | **SI base unit**| **Symbol** |
| Length’          | meter           | m     |
| Mass             | kilogram        | kg    |
| Temperature      | kelvin          | K     |
| Time             | second          | s     |
| Amount of substance | mole       | mol   |
| Luminous intensity | candela   | cd    |
| Electric current | ampere         | A     |

Guide for Reading

- Key Concepts
  - Which five SI base units do chemists commonly use?
  - What metric units are commonly used to measure length, volume, mass, temperature, and energy?

Vocabulary

International System of Units (SI)
- meter (m)
- liter (L)
- kilogram (kg)
- gram (g)
- weight
- temperature
- Celsius scale
- Kelvin scale
- absolute zero
- energy
- joule (J)
- calorie (cal)

Reading Strategy

Summarizing As you read about SI units, summarize the main ideas in the text that follows the red and blue headings.
### Table 3.2

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Meaning</th>
<th>Factor</th>
</tr>
</thead>
<tbody>
<tr>
<td>mega (M)</td>
<td>1 million times larger than the unit it precedes</td>
<td>$10^6$</td>
</tr>
<tr>
<td>kilo (k)</td>
<td>1000 times larger than the unit it precedes</td>
<td>$10^3$</td>
</tr>
<tr>
<td>deci (d)</td>
<td>10 times smaller than the unit it precedes</td>
<td>$10^{-1}$</td>
</tr>
<tr>
<td>centi (c)</td>
<td>100 times smaller than the unit it precedes</td>
<td>$10^{-2}$</td>
</tr>
<tr>
<td>milli (m)</td>
<td>1000 times smaller than the unit it precedes</td>
<td>$10^{-3}$</td>
</tr>
<tr>
<td>micro ($\mu$)</td>
<td>1 million times smaller than the unit it precedes</td>
<td>$10^{-6}$</td>
</tr>
<tr>
<td>nano (n)</td>
<td>1000 million times smaller than the unit it precedes</td>
<td>$10^{-9}$</td>
</tr>
<tr>
<td>pico (p)</td>
<td>1 trillion times smaller than the unit it precedes</td>
<td>$10^{-12}$</td>
</tr>
</tbody>
</table>

### Units and Quantities

As you already know, you don't measure length in kilograms or mass in centimeters. Different quantities require different units. Before you make a measurement, you must be familiar with the units corresponding to the quantity that you are trying to measure.

**Units of Length** Size is an important property of matter. In SI, the basic unit of length, or linear measure, is the meter (m). All measurements of length can be expressed in meters. (The length of a page in this book is about one-fourth of a meter.) For very large and very small lengths, however, it may be more convenient to use a unit of length that has a prefix. Table 3.2 lists the prefixes in common use. For example, the prefix milli- means 1/1000 (one-thousandth), so a millimeter (mm) is 1/1000 of a meter, or 0.001 m. A hyphen (-) measures about 1 mm.

For large distances, it is usually most appropriate to express measurements in kilometers (km). The prefix kilo- means 1000, so 1 km equals 1000 m. A standard marathon race distance of about 42,000 m is more conveniently expressed as 42 km (42 × 1000 m).

**Common metric units of length include the centimeter, meter, and kilometer.** Table 3.3 summarizes the relationships among metric units of length.

### Table 3.3

<table>
<thead>
<tr>
<th>Unit</th>
<th>Relationship</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Kilometer (km)</td>
<td>$1 \text{ km} = 10^3 \text{ m}$</td>
<td>length of about five city blocks $\approx 1 \text{ km}$</td>
</tr>
<tr>
<td>Meter (m)</td>
<td>base unit</td>
<td>height of doorknob from the floor $= 1 \text{ m}$</td>
</tr>
<tr>
<td>Decimeter (dm)</td>
<td>$10^{-1} \text{ dm} = 1 \text{ m}$</td>
<td>diameter of large orange $\approx 1 \text{ dm}$</td>
</tr>
<tr>
<td>Centimeter (cm)</td>
<td>$10^{-2} \text{ cm} = 1 \text{ m}$</td>
<td>width of shirt button $= 1 \text{ cm}$</td>
</tr>
<tr>
<td>Millimeter (mm)</td>
<td>$10^{-3} \text{ mm} = 1 \text{ m}$</td>
<td>thickness of dime $= 1 \text{ mm}$</td>
</tr>
<tr>
<td>Micrometer ($\mu$m)</td>
<td>$10^{-6} \text{ \mu m} = 1 \text{ m}$</td>
<td>diameter of bacterial cell $\approx 1 \text{ \mu m}$</td>
</tr>
<tr>
<td>Nanometer (nm)</td>
<td>$10^{-9} \text{ nm} = 1 \text{ m}$</td>
<td>thickness of RNA molecule $= 1 \text{ nm}$</td>
</tr>
</tbody>
</table>
**Units of Volume** The space occupied by any sample of matter is called its volume. You calculate the volume of any cubic or rectangular solid by multiplying its length by its width by its height. The unit for volume is thus derived from units of length. The SI unit of volume is the amount of space occupied by a cube that is 1 m along each edge. This volume is a cubic meter (m³). An automatic dishwasher has a volume of about 1 m³.

A more convenient unit of volume for everyday use is the liter, a non-SI unit. A liter (L) is the volume of a cube that is 10 centimeters (10 cm) along each edge (10 cm × 10 cm × 10 cm = 1000 cm³ = 1 L). A decimeter (dm) is equal to 10 cm, so 1 L is also equal to 1 cubic decimeter (dm³). A smaller non-SI unit of volume is the milliliter (mL): 1 mL is 1/1000 of a liter. Thus there are 1000 mL in 1 L. Because 1 L is defined as 1000 cm³, 1 mL and 1 cm³ are the same volume. The units milliliter and cubic centimeter are thus used interchangeably. **Common metric units of volume include the liter, milliliter, cubic centimeter, and microliter.** Table 3.4 summarizes the relationships among these units of volume.

There are many devices for measuring liquid volumes, including graduated cylinders, pipets, burets, volumetric flasks, and syringes. Note that the volume of any solid, liquid, or gas will change with temperature (although the change is much more dramatic for gases). Consequently, accurate volume-measuring devices are calibrated at a given temperature—usually 20 degrees Celsius (20°C), which is about normal room temperature.

**Checkpoint** What is the SI unit of volume?

<table>
<thead>
<tr>
<th>Metric Units of Volume</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Unit</strong></td>
</tr>
<tr>
<td>Liter (L)</td>
</tr>
<tr>
<td>Milliliter (mL)</td>
</tr>
<tr>
<td>Cubic centimeter (cm³)</td>
</tr>
<tr>
<td>Microliter (µL)</td>
</tr>
</tbody>
</table>
### Table 3.5

<table>
<thead>
<tr>
<th>Unit</th>
<th>Relationship</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Kilogram (kg) (base unit)</td>
<td>$1 \text{ kg} = 10^3 \text{ g}$</td>
<td>small textbook $= 1 \text{ kg}$</td>
</tr>
<tr>
<td>Gram (g)</td>
<td>$1 \text{ g} = 10^{-3} \text{ kg}$</td>
<td>dollar bill $= 1 \text{ g}$</td>
</tr>
<tr>
<td>Milligram (mg)</td>
<td>$10^3 \text{ mg} = 1 \text{ g}$</td>
<td>ten grains of salt $= 1 \text{ mg}$</td>
</tr>
<tr>
<td>Microgram (µg)</td>
<td>$10^6 \text{ µg} = 1 \text{ g}$</td>
<td>particle of baking powder $= 1 \text{ µg}$</td>
</tr>
</tbody>
</table>

**Units of Mass** The mass of an object is measured in comparison to a standard mass of 1 *kilogram (kg)*, which is the basic SI unit of mass. A kilogram was originally defined as the mass of 1 L of liquid water at 4°C. A cube of water at 4°C measuring 10 cm on each edge would have a volume of 1 L and a mass of 1000 grams (g), or 1 kg. A *gram (g)* is 1/1000 of a kilogram; the mass of 1 cm³ of water at 4°C is 1 g. **Common metric units of mass include the kilogram, gram, milligram, and microgram.** The relationships among units of mass are shown in Table 3.5.

You can use a platform balance to measure the mass of an object. The object is placed on one side of the balance, and standard masses are added to the other side until the balance beam is level. The unknown mass is equal to the sum of the standard masses. Laboratory balances range from very sensitive instruments with a maximum capacity of only a few milligrams to devices for measuring quantities in kilograms. An analytical balance is used to measure objects of less than 100 g and can determine mass to the nearest 0.0001 g (0.1 mg).

The astronaut shown on the surface of the moon in Figure 3.7 weighs one sixth of what he weighs on Earth. The reason for this difference is that the force of gravity on Earth is about six times what it is on the moon. **Weight** is a force that measures the pull on a given mass by gravity. Weight, a measure of force, is different from mass, which is a measure of the quantity of matter. Although the weight of an object can change with its location, its mass remains constant regardless of its location. Objects can thus become weightless, but they can never become massless.

**Figure 3.7** An astronaut’s weight on the moon is one sixth as much as it is on Earth. Earth exerts six times the force of gravity as the moon. **Inferring** How does the astronaut’s mass on the moon compare to his mass on Earth?

How does weight differ from mass?
Units of Temperature  When you hold a glass of hot water, the glass feels hot because heat transfers from the glass to your hand. When you hold an ice cube, it feels cold because heat transfers from your hand to the ice cube. Temperature is a measure of how hot or cold an object is. An object’s temperature determines the direction of heat transfer. When two objects at different temperatures are in contact, heat moves from the object at the higher temperature to the object at the lower temperature.

Almost all substances expand with an increase in temperature and contract as the temperature decreases. (A very important exception is water.) These properties are the basis for the common liquid-in-glass thermometer. The liquid in the thermometer expands and contracts more than the volume of the glass, producing changes in the column height of liquid. Figure 3.8 shows a few different types of thermometers.

Several temperature scales with different units have been devised. Scientists commonly use two equivalent units of temperature, the degree Celsius and the kelvin. The Celsius scale of the metric system is named after the Swedish astronomer Anders Celsius (1701–1744). It uses two readily determined temperatures as reference temperature values: the freezing point and the boiling point of water. The Celsius scale sets the freezing point of water at 0°C and the boiling point of water at 100°C. The distance between these two fixed points is divided into 100 equal intervals, or degrees Celsius (°C).

Another temperature scale used in the physical sciences is the Kelvin, or absolute, scale. This scale is named for Lord Kelvin (1824–1907), a Scottish physicist and mathematician. On the Kelvin scale, the freezing point of water is 273.15 kelvins (K), and the boiling point is 373.15 K. Notice that with the Kelvin scale, the degree sign is not used. Figure 3.9 on the next page compares the Celsius and Kelvin scales. A change of one degree on the Celsius scale is equivalent to one kelvin on the Kelvin scale. The zero point on the Kelvin scale, 0 K, or absolute zero, is equal to −273.15°C. For problems in this text, you can round −273.15°C to −273°C. Because one degree on the Celsius scale is equivalent to one kelvin on the Kelvin scale, converting from one temperature to another is easy. You simply add or subtract 273, as shown in the following equations.

\[ K = °C + 273 \]
\[ °C = K - 273 \]
Figure 3.9 These thermometers show a comparison of the Celsius and Kelvin temperature scales. Note that a 1°C change on the Celsius scale is equal to a 1 K change on the Kelvin scale. 

Interpreting Diagrams What is a change of 10 K equivalent to on the Celsius scale?

SAMPLE PROBLEM 3.4

Converting Between Temperature Scales

Normal human body temperature is 37°C. What is that temperature in kelvins?

1. Analyze List the known and the unknown.

Known
- Temperature in °C = 37°C

Unknown
- Temperature in K = ? K

Use the known value and the equation \( K = °C + 273 \) to calculate the temperature in kelvins.

2. Calculate Solve for the unknown.

Substitute the known value for the Celsius temperature into the equation and solve.

\[
K = °C + 273 \\
= 37 + 273 = 310 K
\]

3. Evaluate Does the result make sense?

You should expect a temperature in this range, since the freezing point of water is 273 K and the boiling point of water is 373 K; normal body temperature is between these two values.

Practice Problems

16. Liquid nitrogen boils at 77.2 K. What is this temperature in degrees Celsius?

17. The element silver melts at 960.8°C and boils at 2212°C. Express these temperatures in kelvins.
Units of Energy  Figure 3.10 shows a house equipped with solar panels. The solar panels convert the radiant energy from the sun into electrical energy that can be used to heat water and power appliances. Energy is the capacity to do work or to produce heat.

Like any other quantity, energy can be measured. The joule and the calorie are common units of energy. The joule (J) is the SI unit of energy. It is named after the English physicist James Prescott Joule (1818–1889). One calorie (cal) is the quantity of heat that raises the temperature of 1 g of pure water by 1°C. Conversions between joules and calories can be carried out using the following relationships.

\[ 1 \text{ J} = 0.2390 \text{ cal} \quad 1 \text{ cal} = 4.184 \text{ J} \]

3.2 Section Assessment

18. Key Concept  Which five SI base units are commonly used in chemistry?

19. Key Concept  Which metric units are commonly used to measure length, volume, mass, temperature, and energy?

20. Name the quantity measured by each of the seven SI base units and give the SI symbol of the unit.

21. What is the symbol and meaning of each prefix?
   a. milli-  
   b. nano-  
   c. deci-  
   d. centi-

22. List the following units in order from largest to smallest: m³, mL, cl, µL, L, dl.

23. What is the volume of a paperback book 21 cm tall, 12 cm wide, and 3.5 cm thick?

24. State the difference between mass and weight.

25. State the relationship between degrees Celsius and kelvins.

26. Surgical instruments may be sterilized by heating at 170°C for 1.5 hr. Convert 170°C to kelvins.

27. State the relationship between joules and calories.

Boiling Points  Look up the boiling points of the first four elements in Group 7A on page R32. Convert these temperatures into kelvins.

Assessment 3.2  Test yourself on the concepts in Section 3.2.
3.3 Conversion Problems

Guide for Reading

Key Concepts
- What happens when a measurement is multiplied by a conversion factor?
- Why is dimensional analysis useful?
- What types of problems are easily solved by using dimensional analysis?

Vocabulary
conversion factor
dimensional analysis

Reading Strategy
Monitoring Your Understanding
Preview the Key Concepts, the section heads, and boldfaced terms. List three things you expect to learn. After reading, state what you learned about each item listed.

Connecting to Your World
Perhaps you have traveled abroad or are planning to do so. If so, you know—or will soon discover—that different countries have different currencies. As a tourist, exchanging money is essential to the enjoyment of your trip. After all, you must pay for your meals, hotel, transportation, gift purchases, and tickets to exhibits and events. Because each country’s currency compares differently with the U.S. dollar, knowing how to convert currency units correctly is very important. Conversion problems are readily solved by a problem-solving approach called dimensional analysis.

Conversion Factors
If you think about any number of everyday situations, you will realize that a quantity can usually be expressed in several different ways. For example, consider the monetary amount $1.

1 dollar = 4 quarters = 10 dimes = 20 nickels = 100 pennies

These are all expressions, or measurements, of the same amount of money. The same thing is true of scientific quantities. For example, consider a distance that measures exactly 1 meter.

1 meter = 10 decimeters = 100 centimeters = 1000 millimeters

These are different ways to express the same length.

Whenever two measurements are equivalent, a ratio of the two measurements will equal 1, or unity. For example, you can divide both sides of the equation 1 m = 100 cm by 1 m or by 100 cm.

\[
\frac{1 \text{ m}}{1 \text{ m}} = \frac{100 \text{ cm}}{1 \text{ m}} = 1 \quad \text{or} \quad \frac{1 \text{ m}}{100 \text{ cm}} = \frac{100 \text{ cm}}{100 \text{ cm}} = 1
\]

A conversion factor is a ratio of equivalent measurements. The ratios 100 cm/1 m and 1 m/100 cm are examples of conversion factors. In a conversion factor, the measurement in the numerator (on the top) is equivalent to the measurement in the denominator (on the bottom). The conversion factors above are read “one hundred centimeters per meter” and “one meter per hundred centimeters.” Figure 3.11 illustrates another way to look at the relationships in a conversion factor. Notice that the smaller number is part of the measurement with the larger unit. That is, a meter is physically larger than a centimeter. The larger number is part of the measurement with the smaller unit.
Conversion factors are useful in solving problems in which a given measurement must be expressed in some other unit of measure. When a measurement is multiplied by a conversion factor, the numerical value is generally changed, but the actual size of the quantity measured remains the same. For example, even though the numbers in the measurements 1 g and 10 dg (decigrams) differ, both measurements represent the same mass. In addition, conversion factors within a system of measurement are defined quantities or exact quantities. Therefore, they have an unlimited number of significant figures, and do not affect the rounding of a calculated answer.

Here are some additional examples of pairs of conversion factors written from equivalent measurements. The relationship between grams and kilograms is 1000 g = 1 kg. The conversion factors are:

\[ \frac{1000 \text{ g}}{1 \text{ kg}} \quad \text{and} \quad \frac{1 \text{ kg}}{1000 \text{ g}} \]

The scale of the micrograph in Figure 3.12 is in nanometers. Using the relationship \(10^9 \text{ nm} = 1 \text{ m}\), you can write the following conversion factors.

\[ \frac{10^9 \text{ nm}}{1 \text{ m}} \quad \text{and} \quad \frac{1 \text{ m}}{10^9 \text{ nm}} \]

Common volumetric units used in chemistry include the liter and the microliter. The relationship \(1 \text{ L} = 10^6 \mu\text{L}\) yields the following conversion factors.

\[ \frac{1 \text{ L}}{10^6 \mu\text{L}} \quad \text{and} \quad \frac{10^6 \mu\text{L}}{1 \text{ L}} \]

Based on what you know about metric prefixes, you should be able to easily write conversion factors that relate equivalent metric quantities.

**Checkpoint** How many significant figures does a conversion factor within a system of measurement have?

**Figure 3.11** The two parts of a conversion factor, the numerator and the denominator, are equal.

**Figure 3.12** In this computer image of atoms, distance is marked off in nanometers (nm). Inferring What conversion factor would you use to convert nanometers to meters?

**Dimensional Analysis**

No single method is best for solving every type of problem. Several good approaches are available, and generally one of the best is dimensional analysis. **Dimensional analysis** is a way to analyze and solve problems using the units, or dimensions, of the measurements. The best way to explain this problem-solving technique is to use it to solve an everyday situation.
SAMPLE PROBLEM 3.5

Using Dimensional Analysis

How many seconds are in a workday that lasts exactly eight hours?

1. **Analyze** List the knowns and the unknown.
   - **Knowns**
     - time worked = 8 h
     - 1 hour = 60 min
     - 1 minute = 60 s
   - **Unknown**
     - seconds worked = ? s

   The first conversion factor must be written with the unit hours in the denominator. The second conversion factor must be written with the unit minutes in the denominator. This will provide the desired unit (seconds) in the answer.

2. **Calculate** Solve for the unknown.

   Start with the known, 8 hours. Use the first relationship (1 hour = 60 minutes) to write a conversion factor that expresses 8 hours as minutes. The unit hours must be in the denominator so that the known unit will cancel. Then use the second conversion factor to change the unit minutes into the unit seconds. This conversion factor must have the unit minutes in the denominator. The two conversion factors can be used together in a simple overall calculation.

   \[
   8 \text{ h} \times \frac{60 \text{ min}}{1 \text{ h}} \times \frac{60 \text{ s}}{1 \text{ min}} = 28,800 \text{ s} \\
   = 2.8800 \times 10^4 \text{ s}
   \]

3. **Evaluate** Does the result make sense?

   The answer has the desired unit (seconds). Since the second is a small unit of time, you should expect a large number of seconds in 8 hours. Before you do the actual arithmetic, it is a good idea to make sure that the units cancel and that the numerator and denominator of each conversion factor are equal to each other. The answer is exact since the given measurement and each of the conversion factors is exact.

**Practice Problems**

- 28. How many minutes are there in exactly one week?
- 29. How many seconds are in exactly a 40-hour work week?

There is usually more than one way to solve a problem. When you first read Sample Problem 3.5, you may have thought about different and equally correct ways to approach and solve the problem. Some problems are easily worked with simple algebra. **Dimensional analysis provides you with an alternative approach to problem solving.** In either case, you should choose the problem-solving method that works best.
**SAMPLE PROBLEM 3.6**

**Using Dimensional Analysis**

The directions for an experiment ask each student to measure 1.84 g of copper (Cu) wire. The only copper wire available is a spool with a mass of 50.0 g. How many students can do the experiment before the copper runs out?

1. **Analyze**  *List the knowns and the unknown.*

   **Knowns**
   - mass of copper available = 50.0 g Cu
   - each student needs 1.84 grams of copper, or \( \frac{1.84 \text{ g Cu}}{\text{student}} \).

   **Unknown**
   - number of students = ?

   From the known mass of copper, calculate the number of students that can do the experiment by using the appropriate conversion factor. The desired conversion is mass of copper \( \longrightarrow \) number of students.

2. **Calculate**  *Solve for the unknown.*

   Because students is the desired unit for the answer, the conversion factor should be written with students in the numerator. Multiply the mass of copper by the conversion factor.

   \[
   50.0 \text{ g Cu} \times \frac{1 \text{ student}}{1.84 \text{ g Cu}} = 27.174 \text{ students} = 27 \text{ students}
   \]

   Note that because students cannot be fractional, the result is shown rounded down to a whole number.

3. **Evaluate**  *Does the result make sense?*

   The unit of the answer (students) is the one desired. The number of students (27) seems to be a reasonable answer. You can make an approximate calculation using the following conversion factor.

   \[
   \frac{1 \text{ student}}{2 \text{ g Cu}}
   \]

   Multiplying the above conversion factor by 50 g Cu gives the approximate answer of 25 students, which is close to the calculated answer.

**Practice Problems**

30. An experiment requires that each student use an 8.5-cm length of magnesium ribbon. How many students can do the experiment if there is a 570-cm length of magnesium ribbon available?

31. A 1.00-degree increase on the Celsius scale is equivalent to a 1.80-degree increase on the Fahrenheit scale. If a temperature increases by 48.0°C, what is the corresponding temperature increase on the Fahrenheit scale?
Converting Between Units

In chemistry, as in many other subjects, you often need to express a measurement in a unit different from the one given or measured initially. Problems in which a measurement with one unit is converted to an equivalent measurement with another unit are easily solved using dimensional analysis.

Suppose that a laboratory experiment requires 7.5 dg of magnesium metal, and 100 students will do the experiment. How many grams of magnesium should your teacher have on hand? Multiplying 100 students by 7.5 dg/student gives you 750 dg. But then you must convert dg to grams. Sample Problem 3.7 shows you how to do the conversion.

SAMPLE PROBLEM 3.7

Converting Between Metric Units

Express 750 dg in grams.

1. Analyze  List the knowns and the unknown.

   Knowns
   • mass = 750 dg
   • 1 g = 10 dg

   Unknown
   • mass = ? g

   The desired conversion is decigrams ——> grams. Using the expression relating the units, 10 dg = 1 g, multiply the given mass by the proper conversion factor.

2. Calculate  Solve for the unknown.

   The correct conversion factor is shown below.

   \[
   \frac{1 \text{ g}}{10 \text{ dg}}
   \]

   Note that the known unit is in the denominator and the unknown unit is in the numerator.

   \[
   750 \text{ dg} \times \frac{1 \text{ g}}{10 \text{ dg}} = 75 \text{ g}
   \]

3. Evaluate  Does the result make sense?

   Because the unit gram represents a larger mass than the unit decigram, it makes sense that the number of grams is less than the given number of decigrams. The unit of the known (dg) cancels, and the answer has the correct unit (g). The answer also has the correct number of significant figures.

Practice Problems

32. Using tables from this chapter, convert the following.
   a. 0.044 km to meters
   b. 4.6 mg to grams
   c. 0.107 g to centigrams

33. Convert the following.
   a. 15 cm³ to liters
   b. 7.38 g to kilograms
   c. 6.7 s to milliseconds
   d. 94.5 g to micrograms
**Multistep Problems** Many complex tasks in your everyday life are best handled by breaking them down into manageable parts. For example, if you were cleaning a car, you might first vacuum the inside, then wash the exterior, then dry the exterior, and finally put on a fresh coat of wax. Similarly, many complex word problems are more easily solved by breaking the solution down into steps.

When converting between units, it is often necessary to use more than one conversion factor. Sample Problem 3.8 illustrates the use of multiple conversion factors.

**Checkpoint** What problem-solving methods can help you solve complex word problems?

---

**SAMPLE PROBLEM 3.8**

**Converting Between Metric Units**

What is 0.073 cm in micrometers?

1. **Analyze** List the knowns and the unknown.
   - **Knowns**
     - length = 0.073 cm = \(7.3 \times 10^{-2}\) cm
     - \(10^2\) cm = 1 m
     - \(1\) m = \(10^6\) \(\mu\)m
   - **Unknown**
     - length = \(?\) \(\mu\)m

   The desired conversion is from centimeters to micrometers. The problem can be solved in a two-step conversion.

2. **Calculate** Solve for the unknown.
   - First change centimeters to meters; then change meters to micrometers: centimeters \(\rightarrow\) meters \(\rightarrow\) micrometers. Each conversion factor is written so that the unit in the denominator cancels the unit in the numerator of the previous factor.
     - \(7.3 \times 10^{-2}\) cm \(\times\) \(\frac{1\ m}{10^2\ cm}\) \(\times\) \(\frac{10^6\ \mu\text{m}}{1\ m}\) = \(7.3 \times 10^2\ \mu\text{m}\)

3. **Evaluate** Does the result make sense?
   - Because a micrometer is a much smaller unit than a centimeter, the answer should be numerically larger than the given measurement. The units have canceled correctly, and the answer has the correct number of significant figures.

---

**Practice Problems**

34. The radius of a potassium atom is 0.227 nm. Express this radius in the unit centimeters.

35. The diameter of Earth is \(1.3 \times 10^4\) km. What is the diameter expressed in decimeters?
**Converting Complex Units** Many common measurements are expressed as a ratio of two units. For example, the results of international car races often give average lap speeds in kilometers per hour. You measure the densities of solids and liquids in grams per cubic centimeter. You measure the gas mileage in a car in miles per gallon of gasoline. If you use dimensional analysis, converting these complex units is just as easy as converting single units. It will just take multiple steps to arrive at an answer.

**SAMPLE PROBLEM 3.9**

**Converting Ratios of Units**

The mass per unit volume of a substance is a property called density. The density of manganese, a metallic element, is $7.21 \text{ g/cm}^3$. What is the density of manganese expressed in units $\text{kg/m}^3$?

1. **Analyze** List the knowns and the unknown.
   
   **Knowns**
   - density of manganese = $7.21 \text{ g/cm}^3$
   - $10^3 \text{ g} = 1 \text{ kg}$
   - $10^6 \text{ cm}^3 = 1 \text{ m}^3$
   
   **Unknown**
   - density manganese = ? $\text{kg/m}^3$

   The desired conversion is $\text{g/cm}^3 \rightarrow \text{kg/m}^3$. The mass unit in the numerator must be changed from grams to kilograms: $\text{g} \rightarrow \text{kg}$. In the denominator, the volume unit must be changed from cubic centimeters to cubic meters: $\text{cm}^3 \rightarrow \text{m}^3$. Note that the relationship between $\text{cm}^3$ and $\text{m}^3$ was determined from the relationship between cm and m. Cubing the relationship $10^2 \text{ cm} = 1 \text{ m}$ yields $(10^2 \text{ cm})^3 = (1 \text{ m})^3$, or $10^6 \text{ cm}^3 = 1 \text{ m}^3$.

2. **Calculate** Solve for the unknown.

   $\frac{7.21 \text{ g}}{1 \text{ cm}^3} \times \frac{1 \text{ kg}}{10^3 \text{ g}} \times \frac{10^6 \text{ cm}^3}{1 \text{ m}^3} = 7.21 \times 10^3 \text{ kg/m}^3$

3. **Evaluate** Does the result make sense?

   Because the physical size of the volume unit $\text{m}^3$ is so much larger than $\text{cm}^3$ ($10^6$ times), the calculated value of the density should be larger than the given value even though the mass unit is also larger ($10^3$ times). The units cancel, the conversion factors are correct, and the answer has the correct ratio of units.

**Practice Problems**

36. Gold has a density of $19.3 \text{ g/cm}^3$. What is the density in kilograms per cubic meter?

37. There are $7.0 \times 10^6$ red blood cells (RBC) in $1.0 \text{ mm}^3$ of blood. How many red blood cells are in $1.0 \text{ L}$ of blood?
Dimensional Analysis

**Purpose**
To apply the problem-solving technique of dimensional analysis to conversion problems.

**Materials**
- 3 inch × 5 inch index cards or paper cut to approximately the same size
- pen

**Procedure**
A conversion factor is a ratio of equivalent measurements. For any relationship, you can write two ratios. On a conversion factor card you can write one ratio on each side of the card.

1. Make a conversion factor card for each metric relationship shown in Tables 3.3, 3.4, and 3.5. Show the inverse of the conversion factor on the back of each card.

2. Use the appropriate conversion factor cards to set up solutions to Sample Problems 3.7 and 3.8. Notice that in each solution, the unit in the denominator of the conversion factor cancels the unit in the numerator of the previous conversion factor.

**Analyze and Conclude**
1. What is the effect of multiplying a given measurement by one or more conversion factors?
2. Use your conversion factor cards to set up solutions to these problems.
   - a. 78.5 cm = ? m
   - b. 0.056 L = ? cm³
   - c. 77 kg = ? mg
   - d. 0.098 nm = ? dm
   - e. 0.96 cm = ? μm
   - f. 0.0067 mm = ? nm

### 3.3 Section Assessment

38. **Key Concept** What happens to the numerical value of a measurement that is multiplied by a conversion factor? What happens to the actual size of the quantity?

39. **Key Concept** Why is dimensional analysis useful?

40. **Key Concept** What types of problems can be solved using dimensional analysis?

41. What conversion factor would you use to convert between these pairs of units?
   - a. minutes to hours
   - b. grams to milligrams
   - c. cubic decimeters to milliliters

42. Make the following conversions. Express your answers in standard exponential form.
   - a. 14.8 g to micrograms
   - b. 3.72 × 10⁻³ kg to grams
   - c. 66.3 L to cubic centimeters

43. An atom of gold has a mass of 3.271 × 10⁻²² g. How many atoms of gold are in 5.00 g of gold?

44. Convert the following. Express your answers in scientific notation.
   - a. 7.5 × 10⁶ J to kilojoules
   - b. 3.9 × 10⁵ mg to decigrams
   - c. 2.21 × 10⁻⁴ dL to microliters

45. Light travels at a speed of 3.00 × 10⁸ cm/s. What is the speed of light in kilometers/hour?
Scale Models

A scale model is a physical or conceptual representation of an object that is proportional in size to the object it represents. Examples include model trains, model airplanes, and dollhouses. Most model trains are built to a scale of 1:87. This ratio means that the model is \( \frac{1}{87} \) the size of an actual train. On the model, 1 cm represents 87 cm on the train.

Scale models aren’t just for hobbyists—scientists and engineers use them, too. A simple scientific model in the classroom is a globe, which is a small-scale model of Earth. (A globe with a diameter of 30 cm has a scale of 1:42,500,000.) Applying Concepts How do you use the scale of a model as a conversion factor?

Computer modeling
By testing a model, engineers can make the product better before it is built. Engineers often design scale models on computers. These automotive engineers are using a computer-aided design (CAD) program to view a digital scale model of a car. Physical models of the car’s wheels are on the desk.

Model building
Architects use both two-dimensional and three-dimensional scale models to design buildings. A common scale for floor plans is 1:48.
Have you ever wondered why some objects float in water, while others sink? If you think that these lily pads float because they are lightweight, you are only partially correct. The ratio of the mass of an object to its volume can be used to determine whether an object floats or sinks in water. For pure water at 4°C, this ratio is 1.000 g/cm³. If an object has a mass-to-volume ratio less than 1.000 g/cm³, it will float in water. If an object has a mass-to-volume ratio greater than this value, it will sink in water.

**Determining Density**

Perhaps someone has tricked you with this question: “Which is heavier, a pound of lead or a pound of feathers?” Most people would not give the question much thought and would incorrectly answer “lead.” Of course, a pound of lead has the same mass as a pound of feathers. What concept, instead of mass, are people really thinking of when they answer this question?

Most people are incorrectly applying a perfectly correct idea: namely, that if a piece of lead and a feather of the same volume are weighed, the lead would have a greater mass than the feather. It would take a much larger volume of feathers to equal the mass of a given volume of lead.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Volume (cm³)</th>
<th>Density (g/cm³)</th>
</tr>
</thead>
<tbody>
<tr>
<td>10 g of Lithium</td>
<td>19</td>
<td>0.53</td>
</tr>
<tr>
<td>10 g of Water</td>
<td>10</td>
<td>1.00</td>
</tr>
<tr>
<td>10 g of Lead</td>
<td>0.88</td>
<td>11.40</td>
</tr>
</tbody>
</table>

**Figure 3.13** A 10-g sample of pure water has less volume than 10 g of lithium, but more volume than 10 g of lead. The faces of the cubes are shown actual size. **Inferring** Which substance has the highest ratio of mass to volume?
<table>
<thead>
<tr>
<th>Solids and Liquids</th>
<th>Gases</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Material</strong></td>
<td><strong>Material</strong></td>
</tr>
<tr>
<td>Gold</td>
<td>Chlorine</td>
</tr>
<tr>
<td>Mercury</td>
<td>Carbon dioxide</td>
</tr>
<tr>
<td>Lead</td>
<td>Argon</td>
</tr>
<tr>
<td>Aluminum</td>
<td>Oxygen</td>
</tr>
<tr>
<td>Table sugar</td>
<td>Air</td>
</tr>
<tr>
<td>Corn syrup</td>
<td>Nitrogen</td>
</tr>
<tr>
<td>Water (4°C)</td>
<td>Neon</td>
</tr>
<tr>
<td>Corn oil</td>
<td>Ammonia</td>
</tr>
<tr>
<td>Ice (0°C)</td>
<td>Methane</td>
</tr>
<tr>
<td>Ethanol</td>
<td>Helium</td>
</tr>
<tr>
<td>Gasoline</td>
<td>Hydrogen</td>
</tr>
</tbody>
</table>

The important relationship in this case is between the object’s mass and its volume. This relationship is called density. **Density** is the ratio of the mass of an object to its volume.

![Density formula](https://via.placeholder.com/150)

A 10.0-cm³ piece of lead, for example, has a mass of 114 g. What, then, is the density of lead? You can calculate it by substituting the mass and volume into the equation above.

\[
\text{Density} = \frac{\text{mass}}{\text{volume}}
\]

\[
\frac{114 \text{ g}}{10.0 \text{ cm}^3} = 11.4 \text{ g/cm}^3
\]

Note that when mass is measured in grams, and volume in cubic centimeters, density has units of grams per cubic centimeter (g/cm³).

Figure 3.13 on page 89 compares the density of three substances. Why does each 10-g sample have a different volume? The volumes vary because the substances have different densities. **Density is an intensive property that depends only on the composition of a substance, not on the size of the sample.** With a mixture, density can vary because the composition of a mixture can vary.

What do you think will happen if corn oil is poured into a glass containing corn syrup? Using Table 3.6, you can see that the density of corn oil is less than the density of corn syrup. For that reason, the oil floats on top of the syrup, as shown in Figure 3.14.

You have probably seen a helium-filled balloon rapidly rise to the ceiling when it is released. Whether a gas-filled balloon will sink or rise when released depends on how the density of the gas compares with the density of air. Helium is less dense than air, so a helium-filled balloon rises. The densities of various gases are listed in Table 3.6.

**Checkpoint** What quantities do you need to measure in order to calculate the density of an object?
Density and Temperature
Experiments show that the volume of most substances increases as the temperature increases. Meanwhile, the mass remains the same despite the temperature and volume changes. Remember that density is the ratio of an object’s mass to its volume. So if the volume changes with temperature (while the mass remains constant), then the density must also change with temperature. The density of a substance generally decreases as its temperature increases. As you will learn in Chapter 15, water is an important exception. Over a certain range of temperatures, the volume of water increases as its temperature decreases. Ice, or solid water, floats because it is less dense than liquid water.

SAMPLE PROBLEM 3.10
Calculating Density
A copper penny has a mass of 3.1 g and a volume of 0.35 cm³. What is the density of copper?

1 Analyze List the knowns and the unknown.
Knobs Unknown
• mass = 3.1 g • density = ? g/cm³
• volume = 0.35 cm³

Use the known values and the following definition of density.

\[
\text{Density} = \frac{\text{mass}}{\text{volume}}
\]

2 Calculate Solve for the unknown.
The equation is already set up to solve for the unknown. Substitute the known values for mass and volume, and calculate the density.

\[
density = \frac{3.1 \text{ g}}{0.35 \text{ cm}^3} = 8.8571 \text{ g/cm}^3
\]

= 8.9 g/cm³ (rounded to two significant figures)

3 Evaluate Does the result make sense?
A piece of copper with a volume of about 0.3 cm³ of copper has a mass of about 3 grams. Thus, about three times that volume of copper, 1 cm³, should have a mass three times larger, about 9 grams. This estimate agrees with the calculated result.

Practice Problems

46. A student finds a shiny piece of metal that she thinks is aluminum. In the lab, she determines that the metal has a volume of 245 cm³ and a mass of 612 g. Calculate the density. Is the metal aluminum?

47. A bar of silver has a mass of 68.0 g and a volume 6.48 cm³. What is the density of silver?
SAMPLE PROBLEM 3.11

Using Density to Calculate Volume

What is the volume of a pure silver coin that has a mass of 14 g? The density of silver (Ag) is 10.5 g/cm³.

1. Analyze List the knowns and the unknown.
   - Knowns
     - mass of coin = 14 g
     - density of silver = 10.5 g/cm³
   - Unknown
     - volume of coin = ? cm³

   You can solve this problem by using density as a conversion factor. You need to convert the mass of the coin into a corresponding volume. The density gives the following relationship between volume and mass.

   \[ 1 \text{ cm}^3 \text{ Ag} = 10.5 \text{ g Ag} \]

   Based on this relationship, you can write the following conversion factor.

   \[ 1 \text{ cm}^3 \text{ Ag} \]
   \[ \frac{10.5 \text{ g Ag}}{10.5 \text{ g Ag}} \]

   Notice that the known unit is in the denominator and the unknown unit is in the numerator.

2. Calculate Solve for the unknown.
   Multiply the mass of the coin by the conversion factor to yield an answer in cm³.

   \[ 14 \text{ g Ag} \times \frac{1 \text{ cm}^3 \text{ Ag}}{10.5 \text{ g Ag}} = 1.3 \text{ cm}^3 \text{ Ag} \]

3. Evaluate Does the result make sense?
   Because a mass of 10.5 g of silver has a volume of 1 cm³, it makes sense that 14.0 g of silver should have a volume slightly larger than 1 cm³. The answer has two significant figures because the given mass has two significant figures.

Practice Problems

48. Use dimensional analysis and the given densities to make the following conversions.
   a. 14.8 g of boron to cm³ of boron. The density of boron is 2.34 g/cm³.
   b. 4.62 g of mercury to cm³ of mercury. The density of mercury is 13.5 g/cm³.

49. Rework the preceding problems by applying the following equation.

   \[ \text{Density} = \frac{\text{mass}}{\text{volume}} \]
Analytical Chemist

Analytical chemists focus on making quantitative measurements. They must be familiar with many analytical techniques to work successfully on a wide variety of tasks. As an analytical chemist, you would spend your time making measurements and calculations to solve laboratory and math-based research problems. You could, for example, be involved in analyzing the composition of biomolecules. Pharmaceutical companies need people to analyze the composition of medicines and research new combinations of compounds to use as drugs. As an analytical chemist, you must be able to think creatively and develop new means for finding solutions.

Many exciting new fields, such as biomedicine and biochemistry, are now hiring analytical chemists. More traditional areas, including industrial manufacturers, also employ analytical chemists. The educational background you need to enter this field is quite extensive. You would need advanced chemical training, including organic chemistry and quantitative chemistry, as well as some training in molecular biology and computer operation. A master’s degree in chemistry may be required, and certain positions require a Ph.D.

3.4 Section Assessment

50. **Key Concept** What determines the density of an object?

51. **Key Concept** How does density vary with temperature?

52. A weather balloon is inflated to a volume of $2.2 \times 10^3$ L with 37.4 g of helium. What is the density of helium in grams per liter?

53. A 68-g bar of gold is cut into 3 equal pieces. How does the density of each piece compare to the density of the original gold bar?

54. A plastic ball with a volume of 19.7 cm$^3$ has a mass of 15.8 g. Would this ball sink or float in a container of gasoline?

55. What is the volume, in cubic centimeters, of a sample of cough syrup that has a mass of 50.0 g? The density of cough syrup is 0.950 g/cm$^3$.

56. What is the mass, in kilograms, of 14.0 L of gasoline? (Assume that the density of gasoline is 0.680 g/cm$^3$.)

**Elements Handbook**

**Density** Look up the densities of the elements in Group 1A on page 96. Which Group 1A elements are less dense than pure water at 4°C?

**Interactive Textbook**

**Assessment 3.4** Test yourself on the concepts in Section 3.4.
Now What Do I Do?

**Purpose**
To solve problems by making accurate measurements and applying mathematics.

**Materials**
- pencil
- paper
- meter stick
- balance
- pair of dice
- aluminum can
- calculator
- small-scale pipet
- water
- a pre- and post-1982 penny
- 8-well strip
- plastic cup

**Procedure**
1. Determine the mass, in grams, of one drop of water. To do this, measure the mass of an empty cup. Add 50 drops of water from a small-scale pipet to the cup and measure its mass again. Subtract the mass of the empty cup from the mass of the cup with water in it. To determine the average mass in grams of a single drop, divide the mass of the water by the number of drops (50). Repeat this experiment until your results are consistent.

2. Determine the mass of a pre-1982 penny and a post-1982 penny.

**Analysis**
Using your experimental data, record the answers to the following questions.

1. What is the average mass of a single drop of water in milligrams? (1 g = 1000 mg)

2. The density of water is 1.00 g/cm³. Calculate the volume of a single drop in cm³ and mL. (1 mL = 1 cm³) What is the volume of a drop in microliters (μL)? (1000 μL = 1 mL)

3. What is the density of water in units of mg/cm³ and mg/mL? (1 g = 1000 mg)

4. Pennies made before 1982 consist of 95.0% copper and 5.0% zinc. Calculate the mass of copper and the mass of zinc in the pre-1982 penny.

5. Pennies made after 1982 are made of zinc with a thin copper coating. They are 97.6% zinc and 2.4% copper. Calculate the mass of copper and the mass of zinc in the newer penny.

6. Why does one penny have less mass than the other?

**You’re the Chemist**
The following small-scale activities allow you to develop your own procedures and analyze the results.

1. **Design It!** Design an experiment to determine if the size of drops varies with the angle at which they are delivered from the pipet. Try vertical (90°), horizontal (0°), and halfway between (45°). Repeat until your results are consistent.

2. **Analyze It!** What is the best angle to hold a pipet for ease of use and consistency of measurement? Explain. Why is it important to expel the air bubbles before you begin the experiment?

3. **Design It!** Make the necessary measurements to determine the volume of aluminum used to make an aluminum soda can. *Hint:* Look up the density of aluminum in your textbook.

4. **Design It!** Design and carry out some experiments to determine the volume of liquid that an aluminum soda can will hold.

5. **Design It!** Measure a room and calculate the volume of air it contains. Estimate the percent error associated with not taking into account the furniture in the room.

6. **Design It!** Make the necessary measurements and do the necessary calculations to determine the volume of a pair of dice. First, ignore the volume of the dots on each face, and then account for the volume of the dots. What is your error and percent error when you ignore the holes?

7. **Design It!** Design an experiment to determine the volume of your body. Write down what measurements you would need to make and what calculations you would do. What additional information might be helpful?
Study Guide

Key Concepts

3.1 Measurements and Their Uncertainty
- Measurements are fundamental to the experimental sciences.
- To evaluate accuracy, the measured value must be compared to the correct value. To evaluate precision, you must compare the values of repeated measurements.
- Calculated answers often depend on the number of significant figures in the values used in the calculation.
- In general, a calculated answer cannot be more precise than the least precise measurement from which it was calculated.

3.2 The International System of Units
- Five commonly used SI base units are the meter, kilogram, kelvin, second, and mole.
- Common metric units of length: cm, m, km.

3.3 Conversion Problems
- Multiplying by a conversion factor does not change the actual size of a measurement.
- Dimensional analysis provides an alternative approach to problem solving.
- Conversion problems are easily solved using dimensional analysis.

3.4 Density
- Density is an intensive property that depends only on the composition of a substance.
- The density of a substance generally decreases as its temperature increases.

Vocabulary
- absolute zero (p. 77)
- accepted value (p. 65)
- accuracy (p. 64)
- calorie (cal) (p. 79)
- Celsius scale (p. 77)
- conversion factor (p. 80)
- density (p. 90)
- dimensional analysis (p. 81)
- energy (p. 79)
- error (p. 65)
- experimental value (p. 65)
- gram (g) (p. 76)
- International System of Units (SI) (p. 73)
- joule (J) (p. 79)
- Kelvin scale (p. 77)
- kilogram (kg) (p. 76)
- liter (L) (p. 75)
- measurement (p. 63)
- meter (m) (p. 74)
- percent error (p. 65)
- precision (p. 64)
- scientific notation (p. 63)
- significant figures (p. 66)
- temperature (p. 77)
- weight (p. 76)

Key Equations
- Error = experimental value – accepted value
- Percent error = \( \frac{|error|}{accepted\ value} \times 100\% \)
- \( K = °C + 273 \) and \( °C = K - 273 \)
- \( 1 \text{ J} = 0.2390 \text{ cal} \) and \( 1 \text{ cal} = 4.184 \text{ J} \)
- Density = \( \frac{mass}{volume} \)

Organizing Information
Use these terms to construct a concept map that organizes the major ideas of this chapter.

Interactive Textbook
Concept Map 3 Solve the concept map with the help of an interactive guided tutorial.

with ChemASAP
3.1 Measurements and Their Uncertainty

57. Three students made multiple weighings of a copper cylinder, each using a different balance. Describe the accuracy and precision of each student's measurements if the correct mass of the cylinder is 47.32 g.

<table>
<thead>
<tr>
<th>Mass of Cylinder (g)</th>
<th>Lissa</th>
<th>Lamont</th>
<th>Leigh Anne</th>
</tr>
</thead>
<tbody>
<tr>
<td>Weighing 1</td>
<td>47.13</td>
<td>47.45</td>
<td>47.95</td>
</tr>
<tr>
<td>Weighing 2</td>
<td>47.94</td>
<td>47.39</td>
<td>47.91</td>
</tr>
<tr>
<td>Weighing 3</td>
<td>46.83</td>
<td>47.42</td>
<td>47.89</td>
</tr>
<tr>
<td>Weighing 4</td>
<td>47.47</td>
<td>47.41</td>
<td>47.93</td>
</tr>
</tbody>
</table>

58. How many significant figures are in each underlined measurement?
   a. 60 s = 1 min
   b. 47.70 g of copper
   c. 1 km = 1000 m
   d. 25 computers

59. Round off each of these measurements to three significant figures.
   a. 98.473 L
   b. 0.000 763 21 cg
   c. 57.048 m
   d. 12.17°C

60. Round off each of the answers correctly.
   a. $8.7 \, \text{g} + 15.43 \, \text{g} + 19 \, \text{g} = 43.13 \, \text{g}$
   b. $853.2 \, \text{L} - 627.443 \, \text{L} = 225.757 \, \text{L}$
   c. $38.742 \, \text{kg} + 0.421 = 92.023 \, \text{75 kg}$
   d. $5.40 \, \text{m} \times 3.21 \, \text{m} \times 1.871 \, \text{m} = 32.431 \, 914 \, \text{m}^3$

61. Express each of the rounded-off answers in Questions 59 and 60 in scientific notation.

62. How are the error and the percent error of a measurement calculated?

3.2 The International System of Units

63. List the SI base unit of measurement for each of these quantities.
   a. time
   b. length
   c. temperature
   d. mass

64. Order these units from smallest to largest: cm, μm, km, mm, m, nm, dm, pm. Then give each measurement in terms of meters.

65. Measure each of the following dimensions using a unit with the appropriate prefix.
   a. the height of this letter I
   b. the width of Table 3.3
   c. the height of this page

66. The melting point of silver is 962°C. Express this temperature in kelvins.

3.3 Conversion Problems

67. What is the name given to a ratio of two equivalent measurements?

68. What must be true for a ratio of two measurements to be a conversion factor?

69. How do you know which unit of a conversion factor must be in the denominator?

70. Make the following conversions.
   a. 157 cs to seconds
   b. 42.7 L to milliliters
   c. 261 nm to millimeters
   d. 0.065 km to decimeters
   e. 642 cg to kilograms
   f. $8.25 \times 10^2$ cg to nanograms

71. Make the following conversions.
   a. 0.44 mL/min to microliters per second
   b. 7.86 g/cm² to milligrams per square millimeter
   c. 1.54 kg/L to grams per cubic centimeter

72. How many milliliters are contained in 1 m³?

73. Complete this table so that all the measurements in each row have the same value.

<table>
<thead>
<tr>
<th></th>
<th>mg</th>
<th>g</th>
<th>cg</th>
<th>kg</th>
</tr>
</thead>
<tbody>
<tr>
<td>(a)</td>
<td></td>
<td>(b)</td>
<td></td>
<td>(c)</td>
</tr>
<tr>
<td>$6.8 \times 10^3$</td>
<td>(d)</td>
<td>(e)</td>
<td></td>
<td>(f)</td>
</tr>
<tr>
<td>(g)</td>
<td></td>
<td>(h)</td>
<td></td>
<td>(i)</td>
</tr>
</tbody>
</table>

3.4 Density

74. What equation is used to determine the density of an object?

75. Would the density of a person be the same on the surface of Earth and on the surface of the moon? Explain.

76. A shiny, gold-colored bar of metal weighing 57.3 g has a volume of 4.7 cm³. Is the bar of metal pure gold?

77. Three balloons filled with neon, carbon dioxide, and hydrogen are released into the atmosphere. Using the data in Table 3.6 on page 90, describe the movement of each balloon.
Understanding Concepts

78. List two possible reasons for reporting precise, but inaccurate, measurements.

79. Rank these numbers from smallest to largest.
   a. \(5.3 \times 10^4\)  
   b. \(57 \times 10^3\)  
   c. \(4.9 \times 10^{-2}\)  
   d. 0.0057  
   e. \(5.1 \times 10^{-3}\)  
   f. 0.0072 \(\times 10^2\)

80. Comment on the accuracy and precision of these basketball free-throw shooters.
   a. 99 of 100 shots are made.
   b. 99 of 100 shots hit the front of the rim and bounce off.
   c. 33 of 100 shots are made; the rest miss.

81. Fahrenheit is a third temperature scale. Plot the data in the table and use the graph to derive an equation for the relationship between the Fahrenheit and Celsius temperature scales.

<table>
<thead>
<tr>
<th>Example</th>
<th>°C</th>
<th>°F</th>
</tr>
</thead>
<tbody>
<tr>
<td>Melting point of selenium</td>
<td>221</td>
<td>430</td>
</tr>
<tr>
<td>Boiling point of water</td>
<td>100</td>
<td>212</td>
</tr>
<tr>
<td>Normal body temperature</td>
<td>37</td>
<td>98.6</td>
</tr>
<tr>
<td>Freezing point of water</td>
<td>0</td>
<td>32</td>
</tr>
<tr>
<td>Boiling point of chlorine</td>
<td>-34.6</td>
<td>-30.2</td>
</tr>
</tbody>
</table>

82. Which would melt first, germanium with a melting point of 1210 K or gold with a melting point of 1064°C?

83. Write six conversion factors involving these units of measure: 1 g = \(10^2\) cg = \(10^3\) mg.

84. A 2.00-kg sample of bituminous coal is composed of 1.30 kg of carbon, 0.20 kg of ash, 0.15 kg of water, and 0.35 kg of volatile (gas-forming) material. Using this information, determine how many kilograms of carbon are in 125 kg of this coal.

85. A piece of wood sinks in ethanol but floats in gasoline. Give a range of possible densities for the wood.

86. The density of dry air measured at 25°C is \(1.19 \times 10^{-3}\) g/cm³. What is the volume of 50.0 g of air?

87. A flask that can hold 158 g of water at 4°C can hold only 127 g of ethanol at the same temperature. What is the density of ethanol?

88. A watch loses 0.15 s every minute. How many minutes will the watch lose in 1 day?

89. A tank measuring 28.6 cm by 73.0 mm by 0.72 m is filled with olive oil. The oil in the tank has a mass of \(1.38 \times 10^4\) g. What is the density of olive oil in kilograms per liter?

90. Alkanes are a class of molecules that have the general formula \(C_nH_{2n+2}\), where \(n\) is an integer (whole number). The table below gives the boiling points for the first five alkanes with an odd number of carbon atoms. Using the table, construct a graph with number of carbon atoms on the x-axis.

<table>
<thead>
<tr>
<th>Boiling point (°C)</th>
<th>Number of carbon atoms</th>
</tr>
</thead>
<tbody>
<tr>
<td>-162.0</td>
<td>1</td>
</tr>
<tr>
<td>-42.0</td>
<td>3</td>
</tr>
<tr>
<td>36.0</td>
<td>5</td>
</tr>
<tr>
<td>98.0</td>
<td>7</td>
</tr>
<tr>
<td>151.0</td>
<td>9</td>
</tr>
</tbody>
</table>

a. What are the approximate boiling points for the \(C_2\), \(C_4\), \(C_6\), and \(C_8\) alkanes?

b. Which of these nine alkanes are gases at room temperature (20°C)?

c. How many of these nine alkanes are liquids at 350 K?

d. What is the approximate increase in boiling point per additional carbon atom in these alkanes?

91. Earth is approximately \(1.5 \times 10^8\) km from the sun. How many minutes does it take light to travel from the sun to Earth? The speed of light is \(3.0 \times 10^8\) m/s.

92. What is the mass of a cube of aluminum that is 3.0 cm on each edge? The density of aluminum is 2.7 g/cm³.

93. The average density of Earth is 5.52 g/cm³. Express this density in units of kg/dm³.

94. How many kilograms of water (at 4°C) are needed to fill an aquarium that measures 40.0 cm by 20.0 cm by 30.0 cm?
Critical Thinking

95. Is it possible for an object to lose weight but at the same time not lose mass? Explain.

96. One of the first mixtures of metals, called amalgams, used by dentists for tooth fillings consisted of 26.0 g of silver, 10.8 g of tin, 2.4 g of copper, and 0.8 g of zinc. How much silver is in a 25.0 g sample of this amalgam?

97. A cheetah can run 112 km/h over a 100-m distance. What is this speed in meters per second?

98. You are hired to count the number of ducks on three northern lakes during the summer. In the first lake, you estimate 500,000 ducks, in the second 250,000 ducks, and in the third 100,000 ducks. You write down that you have counted 850,000 ducks. As you drive away, you see 15 ducks fly in from the south and land on the third lake. Do you change the number of ducks that you report? Justify your answer.

99. What if ice were more dense than water? It would certainly be easier to pour water from a pitcher of ice cubes and water. Can you imagine situations of more consequence?

100. Why is there a range of values given for the density of gasoline on Table 3.6 on page 90?

101. Plot these data that show how the mass of sulfur increases with an increase in volume. Determine the density of sulfur from the slope of the line.

<table>
<thead>
<tr>
<th>Volume of sulfur (cm³)</th>
<th>Mass of sulfur (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>11.4</td>
<td>23.5</td>
</tr>
<tr>
<td>29.2</td>
<td>60.8</td>
</tr>
<tr>
<td>55.5</td>
<td>115</td>
</tr>
<tr>
<td>81.1</td>
<td>168</td>
</tr>
</tbody>
</table>

102. At 20°C, the density of air is 1.20 g/L. Nitrogen's density is 1.17 g/L. Oxygen's density is 1.33 g/L.
   a. Will balloons filled with oxygen and balloons filled with nitrogen rise or sink in air?
   b. Air is mainly a mixture of nitrogen and oxygen. Which gas is the main component? Explain.

Concept Challenge

103. The mass of a cube of iron is 355 g. Iron has a density of 7.87 g/cm³. What is the mass of a cube of lead that has the same dimensions?

104. Sea water contains $8.0 \times 10^{-1} \text{ cg}$ of the element strontium per kilogram of sea water. Assuming that all the strontium could be recovered, how many grams of strontium could be obtained from one cubic meter of sea water? Assume the density of sea water is 1.0 g/mL.

105. The density of dry air at 20°C is 1.20 g/L. What is the mass of air, in kilograms, of a room that measures 25.0 m by 15.0 m by 4.0 m?

106. Different volumes of the same liquid were added to a flask on a balance. After each addition of liquid, the mass of the flask with the liquid was measured. Graph the data using mass as the dependent variable. Use the graph to answer these questions.

<table>
<thead>
<tr>
<th>Volume (mL)</th>
<th>Mass (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>14</td>
<td>103.0</td>
</tr>
<tr>
<td>27</td>
<td>120.4</td>
</tr>
<tr>
<td>41</td>
<td>139.1</td>
</tr>
<tr>
<td>56</td>
<td>157.9</td>
</tr>
<tr>
<td>82</td>
<td>194.1</td>
</tr>
</tbody>
</table>

   a. What is the mass of the flask?
   b. What is the density of the liquid?

107. A 34.5-g gold nugget is dropped into a graduated cylinder containing water. By how many milliliters does the measured volume increase? The density of water is 1.0 g/mL. The density of gold is 19.3 g/cm³.

108. Equal amounts of mercury, water, and corn oil are added to a beaker.
   a. Describe the arrangement of the layers of liquids in the beaker.
   b. A small sugar cube is added to the beaker. Describe its location.
   c. What change will occur to the sugar cube over time?
Select the choice that best answers each question or completes each statement.

1. Which of these series of units is ordered from smallest to largest?
   a. μg, cg, mg, kg
   b. mm, dm, m, km
   c. μs, ns, cs, s
   d. nL, mL, dL, L

2. Which answer represents the measurement 0.00428 g rounded to two significant figures?
   a. 4.28 \times 10^2 g
   b. 4.3 \times 10^{-2} g
   c. 4.3 \times 10^2 g
   d. 4.0 \times 10^{-2} g

3. An over-the-counter medicine has 325 mg of its active ingredient per tablet. How many grams does this mass represent?
   a. 325,000 g
   b. 32.5 g
   c. 3.25 g
   d. 0.325 g

4. If 10^4 μm = 1 cm, how many μm^3 = 1 cm^3?
   a. 10^6
   b. 10^8
   c. 10^4
   d. 10^{12}

5. How many meters does a car moving at 95 km/h travel in 1.0 s?
   a. 1.6 m
   b. 340 m
   c. 1600 m
   d. 26 m

6. If a substance contracts when it freezes, its
   a. density will remain the same.
   b. density will increase.
   c. density will decrease.
   d. change in density cannot be predicted.

For Questions 7–9, identify the known and the unknown. Include units in your answers.

7. The density of water is 1.0 g/mL. How many deciliters of water will fill a 0.5-L bottle?

8. A clock loses 4 minutes every day. How many seconds does the clock lose in 1 minute?

9. A graduated cylinder contains 44.2 mL of water. A 48.6-g piece of metal is carefully dropped into the cylinder. When the metal is completely covered with water, the water rises to the 51.3-mL mark. What is the density of the metal?

Use the atomic windows below to answer Questions 10 and 11.

![Atomic Windows](image)

The atomic windows represent particles of the same gas occupying the same volume at the same temperature. The systems differ only in the number of gas particles per unit volume.

10. List the windows in order of decreasing density.

11. Compare the density of the gas in window (a) to the density of the gas in window (b).

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For each question there are two statements. Decide whether each statement is true or false. Then decide whether Statement II is a correct explanation for Statement I.

**Statement I**

12. There are five significant figures in the measurement 0.00450 m.

13. Precise measurements will always be accurate measurements.

14. A temperature in kelvins is always numerically larger than the same temperature in degrees Celsius.

**Statement II**

BECAUSE All zeros to the right of a decimal point in a measurement are significant.

BECAUSE A value that is measured 10 times in a row must be accurate.

BECAUSE A temperature in kelvins equals a temperature in degrees Celsius plus 273.