3 Scientific Measurement

INSIDE:

- **3.1** Using and Expressing Measurements
- **3.2** Units of Measurement
- **3.3** Solving Conversion Problems

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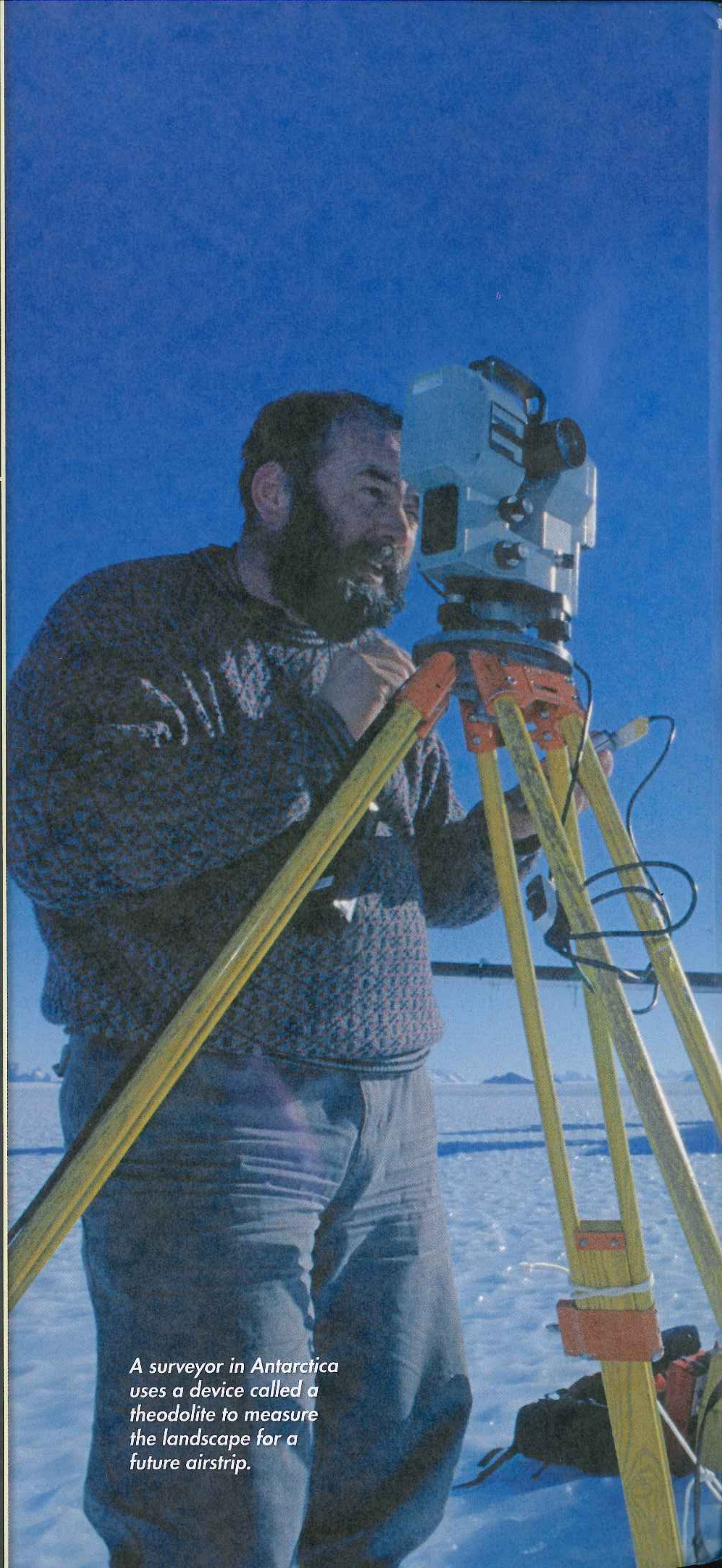
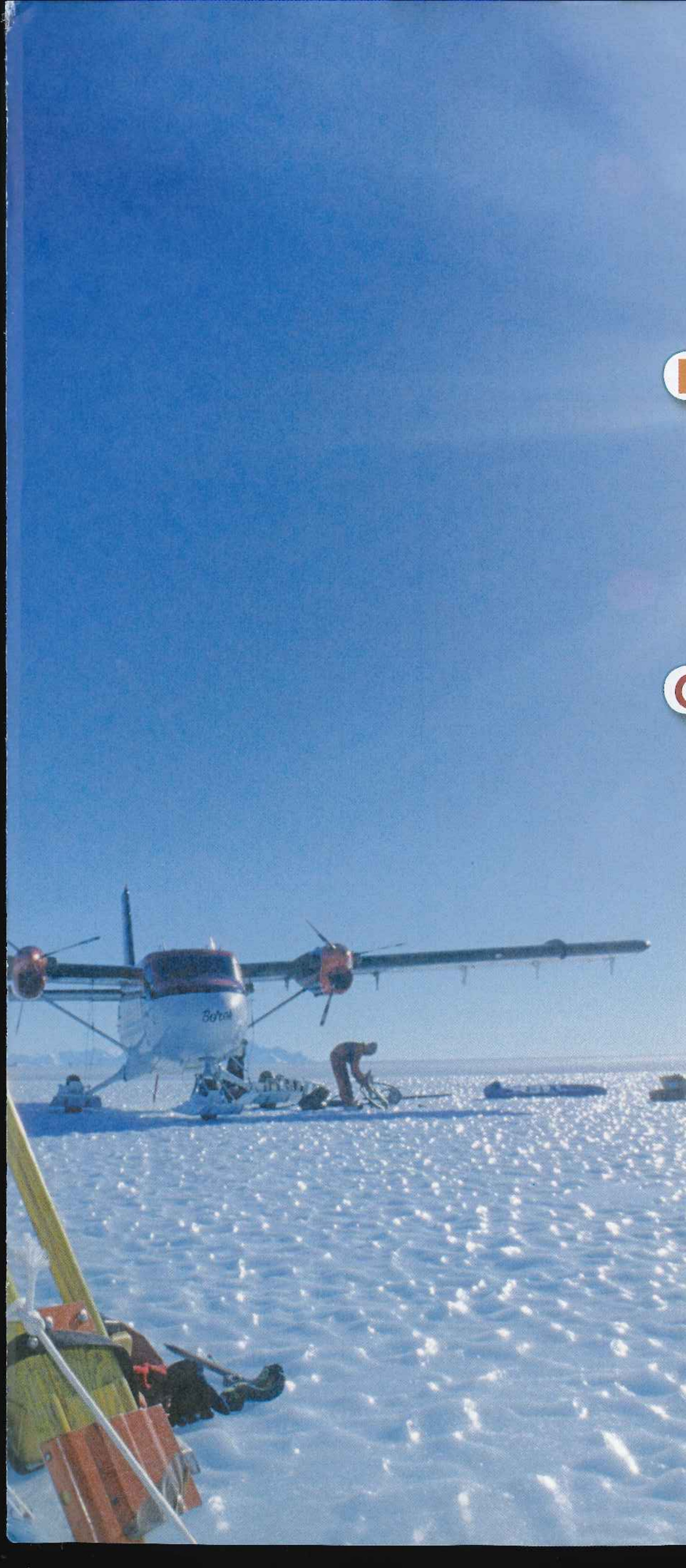


A surveyor in Antarctica uses a device called a theodolite to measure the landscape for a future airstrip.



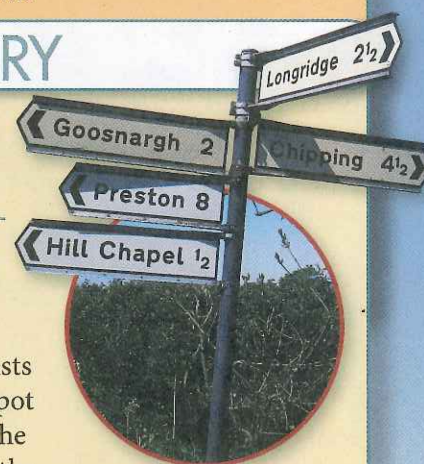
BIGIDEA

QUANTIFYING MATTER

Essential Questions:

1. *How do scientists express the degree of uncertainty in their measurements?*
2. *How is dimensional analysis used to solve problems?*

CHEMYSTERY



Just Give Me a Sign

While traveling in a foreign country, you happen to get lost, as many tourists do. But then you spot these signs along the road. If you know the distance to your destination, you can find your way. However, in the signs shown here, the distances are listed as numbers with no units attached. For example, is Preston 8 kilometers away or 8 miles away? Is there any way to know for sure?

► Connect to the **BIG**IDEA As you read the chapter, try to familiarize yourself with common metric units used in science.

NATIONAL SCIENCE EDUCATION STANDARDS

A-1, E-2

3.1 Using and Expressing Measurements

CHEMISTRY & YOU

Q: *How do you measure a photo finish?* You probably know that a 100-meter dash is timed in seconds. But if it's a close finish, measuring each runner's time to the nearest second will not tell you who won. That's why sprint times are often measured to the nearest hundredth of a second (0.01 s). Chemistry also requires making accurate and often very small measurements.

Key Questions

How do you write numbers in scientific notation?

How do you evaluate accuracy and precision?

Why must measurements be reported to the correct number of significant figures?

Vocabulary

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

Scientific Notation

How do you write numbers in scientific notation?

Everyone makes and uses measurements. A **measurement** is a quantity that has both a number and a unit. Your height (66 inches), your age (15 years), and your body temperature (37°C) are examples of measurements.

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. 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 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 cumbersome. You can work more easily with these numbers by writing them in scientific 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 can be written in scientific notation as 6.02×10^{23}. The coefficient in this number is 6.02. The power of 10, or exponent, is 23. **In scientific notation, the coefficient is always a number greater than or equal to one and less than ten. The exponent is an integer.** A positive exponent indicates how many times the coefficient must be multiplied by 10. A negative exponent indicates how many times the coefficient must be divided by 10. Figure 3.1 shows a magnified view of a human hair, which has a diameter of about 0.00007 m, or 7×10^{-5} m.

When writing numbers greater than ten in scientific notation, the exponent is positive and equals the number of places that the original decimal point has been moved to the left.

$$6{,}300{,}000. = 6.3 \times 10^{6} \qquad 94{,}700. = 9.47 \times 10^{4}$$

Numbers less than one have a negative exponent when written in scientific notation. The value of the exponent equals the number of places the decimal has been moved to the right.

$$0.000\,008 = 8 \times 10^{-6} \qquad 0.00736 = 7.36 \times 10^{-3}$$

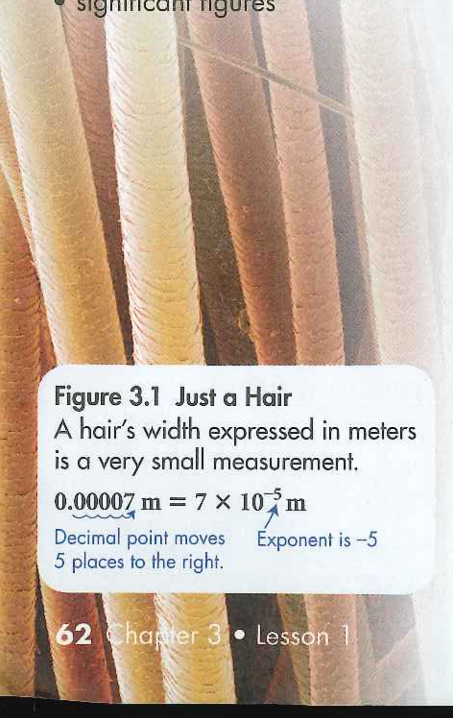


Figure 3.1 Just a Hair
A hair's width expressed in meters is a very small measurement.

0.00007 m $= 7 \times 10^{-5}$ m

Decimal point moves 5 places to the right. Exponent is −5

Multiplication and Division To multiply numbers written in scientific notation, multiply the coefficients and add the exponents.

$$(3 \times 10^4) \times (2 \times 10^2) = (3 \times 2) \times 10^{4+2} = 6 \times 10^6$$

$$(2.1 \times 10^3) \times (4.0 \times 10^{-7}) = (2.1 \times 4.0) \times 10^{3+(-7)} = 8.4 \times 10^{-4}$$

To divide numbers written in scientific notation, divide the coefficients and subtract the exponent in the denominator from the exponent in the numerator.

$$\frac{3.0 \times 10^5}{6.0 \times 10^2} = \left(\frac{3.0}{6.0}\right) \times 10^{5-2} = 0.5 \times 10^3 = 5.0 \times 10^2$$

Addition and Subtraction If you want to add or subtract numbers expressed in scientific notation and you are not using a calculator, then the exponents must be the same. In other words, the decimal points must be aligned before you add or subtract the numbers. For example, when adding 5.4×10^3 and 8.0×10^2, first rewrite the second number so that the exponent is a 3. Then add the numbers.

$$\begin{aligned}(5.4 \times 10^3) + (8.0 \times 10^2) &= (5.4 \times 10^3) + (0.80 \times 10^3)\\ &= (5.4 + 0.80) \times 10^3\\ &= 6.2 \times 10^3\end{aligned}$$

Sample Problem 3.1

Using Scientific Notation

Solve each problem and express the answer in scientific notation.

a. $(8.0 \times 10^{-2}) \times (7.0 \times 10^{-5})$ **b.** $(7.1 \times 10^{-2}) + (5 \times 10^{-3})$

1 Analyze **Identify the relevant concepts.** To multiply numbers in scientific notation, multiply the coefficients and add the exponents. To add numbers in scientific notation, the exponents must match. If they do not, then adjust the notation of one of the numbers.

2 Solve **Apply the concepts to this problem.**

Multiply the coefficients and add the exponents.

$$\begin{aligned}\textbf{a. } (8.0 \times 10^{-2}) \times (7.0 \times 10^{-5}) &= (8.0 \times 7.0) \times 10^{-2+(-5)}\\ &= 56 \times 10^{-7}\\ &= 5.6 \times 10^{-6}\end{aligned}$$

Rewrite one of the numbers so that the exponents match. Then add the coefficients.

$$\begin{aligned}\textbf{b. } (7.1 \times 10^{-2}) + (5 \times 10^{-3}) &= (7.1 \times 10^{-2}) + (0.5 \times 10^{-2})\\ &= (7.1 + 0.5) \times 10^{-2}\\ &= 7.6 \times 10^{-2}\end{aligned}$$

1. Solve each problem and express the answer in scientific notation.

a. $(6.6 \times 10^{-8}) + (5.0 \times 10^{-9})$

b. $(9.4 \times 10^{-2}) - (2.1 \times 10^{-2})$

2. Calculate the following and write your answer in scientific notation:

$$\frac{6.6 \times 10^6}{(8.8 \times 10^{-2}) \times (2.5 \times 10^3)}$$

Accuracy, Precision, and Error

How do you evaluate accuracy and precision?

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, irrespective of the actual value. **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 in Figure 3.2 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.

Figure 3.2 Accuracy vs. Precision
The distribution of darts illustrates the difference between accuracy and precision.
Use Analogies ***Which outcome describes a scenario in which you properly measure an object's mass three times using a balance that has not been zeroed?***

Good Accuracy, Good Precision
Closeness to the bull's-eye indicates a high degree of accuracy. The closeness of the darts to one another indicates high precision.

Poor Accuracy, Good Precision
Precision is high because of the closeness of grouping—thus, the high level of reproducibility. But the results are inaccurate.

Poor Accuracy, Poor Precision
The darts land far from one another and from the bull's-eye. The results are both inaccurate and imprecise.

Determining Error 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 at these conditions is actually 100.0°C.

There is a difference between the **accepted value,** which is the correct value for the measurement 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** of a measurement is the absolute value of the error divided by the accepted value, multiplied by 100%.

$$\text{Percent error} = \frac{|\text{error}|}{\text{accepted value}} \times 100\%$$

READING SUPPORT

Build Reading Skills: ***Inference*** As you read, try to identify some of the factors that cause experimental error. ***What factors might result in inaccurate measurements? What factors might result in imprecise measurements?***

Sample Problem 3.2

Calculating Percent Error

The boiling point of pure water is measured to be 99.1°C. Calculate the percent error.

1 Analyze List the knowns and unknown. The accepted value for the boiling point of pure water is 100°C. Use the equations for error and percent error to solve the problem.

KNOWNS	UNKNOWN
Experimental value = 99.1°C	Percent error = ?
Accepted value = 100.0°C	

2 Calculate Solve for the unknown.

Start with the equation for percent error.

$$\text{Percent error} = \frac{|\text{error}|}{\text{accepted value}} \times 100\%$$

Substitute the equation for error, and then plug in the known values.

$$\text{Percent error} = \frac{|\text{experimental value} - \text{accepted value}|}{\text{accepted value}} \times 100\%$$

$$= \frac{|99.1°C - 100.0°C|}{100.0°C} \times 100\%$$

$$= \frac{0.9\cancel{°C}}{100.0\cancel{°C}} \times 100\% = 0.9\%$$

3 Evaluate Does the result make sense? The experimental value was off by about 1°C, or $\frac{1}{100}$ of the accepted value (100°C). The answer makes sense.

3. A student measures the depth of a swimming pool to be 2.04 meters at its deepest end. The accepted value is 2.00 m. What is the student's percent error?

Think about it: Using the absolute value of the error means that percent error will always be a positive value.

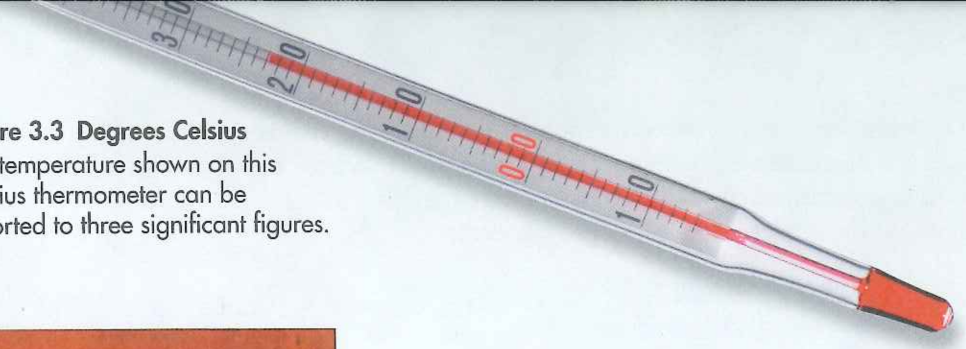


Figure 3.3 Degrees Celsius The temperature shown on this Celsius thermometer can be reported to three significant figures.

Significant Figures

Why must measurements be reported to the correct number of significant figures?

Look at the reading of the thermometer shown in Figure 3.3. If you use a liquid-filled thermometer that is calibrated in 1°C intervals, you can easily read the temperature to the nearest degree. With the same thermometer, however, you can also estimate the temperature to about the nearest tenth of a degree by noting the closeness of the liquid inside to the calibrations. Looking at Figure 3.3, suppose you estimate that the temperature lies between 22°C and 23°C, at 22.9°C. This estimated number has three digits. The first two digits (2 and 2) are known with certainty. But the rightmost digit (9) has been estimated and involves some uncertainty. These 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 3.4 can be used to make successively more precise measurements.

More on precision in measurements *online.* KINETIC ART

Figure 3.4 Increasing Precision Three differently calibrated meter sticks are used to measure a door's width. A meter stick calibrated in 0.1-m (1 dm) intervals is more precise than one calibrated in a 1-m interval but less precise than one calibrated in 0.01-m (1 cm) intervals. **Measure** *How many significant figures are reported in each measurement?*

Determining Significant Figures in Measurements To determine whether a digit in a measured value is significant, you need to apply the following rules.

1. Every nonzero digit in a reported measurement is assumed to be significant.

24.7 meters
0.743 meter
714 meters

Each of these measurements has three significant figures.

---

2. Zeros appearing between nonzero digits are significant.

7003 meters
40.79 meters
1.503 meters

Each of these measurements has four significant figures.

---

3. Leftmost zeros appearing in front of nonzero digits are not significant. They act as placeholders. By writing the measurements in scientific notation, you can eliminate such placeholding zeros.

0.0071 meter = 7.1×10^{-3} meter
0.42 meter = 4.2×10^{-1} meter
0.000099 meter = 9.9×10^{-5} meter

Each of these measurements has only two significant figures.

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4. Zeros at the end of a number and to the right of a decimal point are always significant.

43.00 meters
1.010 meters
9.000 meters

Each of these measurements has four significant figures.

---

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.

300 meters (one significant figure)
7000 meters (one significant figure)
27,210 meters (four significant figures)

The zeros in these measurements are not significant.

If such zeros were known measured values, however, then they would be significant. Writing the value in scientific notation makes it clear that these zeros are significant.

300 meters = 3.00×10^{2} meters (three significant figures)

The zeros in this measurement are significant.

---

6. There are two situations in which numbers have an unlimited number of significant figures. The first involves counting. A number that is counted is exact.

23 people in your classroom

This measurement is a counted value, so it has an unlimited number of significant figures.

The second situation involves exactly defined quantities such as those found within a system of measurement.

60 min = 1 hr
100 cm = 1 m

Each of these numbers has an unlimited number of significant figures.

Sample Problem 3.3

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

Make sure you understand the rules for counting significant figures (on the previous page) before you begin, okay?

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. These locations are known by inspecting each measurement value.

2 Solve **Apply the concepts to this problem.**

Apply the rules for determining significant figures. 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)

4. Count the significant figures in each measured length.
- **a.** 0.057 30 meter
- **b.** 8765 meters
- **c.** 0.000 73 meter
- **d.** 40.007 meters

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

CHEMISTRY & YOU

Q: *Suppose that the winner of a 100-meter dash finishes the race in 9.98 seconds. The runner in second place has a time of 10.05 seconds. How many significant figures are in each measurement? Is one measurement more accurate than the other? Explain your answer.*

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. However, each of the measurements used in the calculation is expressed to only two significant figures. As a result, the answer must also be reported to two significant figures (42 m^2). 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.

SampleProblem 3.4

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.** Using the rules for determining significant figures, round the number in each measurement. Then apply the rules for expressing numbers in scientific notation.

2 Solve **Apply the concepts to this problem.**

Starting from the left, count the first four digits that are significant. 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 = **3.147×10^2 meters**

Starting from the left, count the first two digits that are significant. The arrow points to the digit immediately following the second significant digit.

b. 0.001 775 meters
↑
7 is greater than 5, so round up.
0.0018 meter = **1.8×10^{-3} meter**

Starting from the left, count the first two digits that are significant. The arrow points to the digit immediately following the second significant digit.

c. 8792 meters
↑
9 is greater than 5, so round up.
8800 meters = **8.8×10^3 meters**

6. Round each measurement to three significant figures. Write your answers in scientific notation.

- **a.** 87.073 meters
- **b.** 4.3621×10^8 meters
- **c.** 0.01552 meter
- **d.** 9009 meters
- **e.** 1.7777×10^{-3} meter
- **f.** 629.55 meters

7. Round each measurement in Problem 6 to one significant figure. Write each of your answers in scientific notation.

If you're already familiar with rounding numbers, you can skip to Sample Problems 3.5 and 3.6.

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. Sample Problem 3.5 gives examples of rounding in addition and subtraction.

Multiplication and Division In calculations involving multiplication and division (such as those in Sample Problem 3.6), 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.

Sample Problem 3.5

Significant Figures in Addition and Subtraction

Perform the following addition and subtraction operations. Give each answer to the correct number of significant figures.

a. 12.52 meters + 349.0 meters + 8.24 meters

b. 74.626 meters − 28.34 meters

1 Analyze **Identify the relevant concepts.** Perform the specified math operation, and then round the answer to match the measurement with the least number of decimal places.

2 Solve **Apply the concepts to this problem.**

Align the decimal points and add the numbers.

a.

$$\begin{array}{rl} 12.52 & \text{meters} \\ 349.0 & \text{meters} \\ +\ \ 8.24 & \text{meters} \\ \hline 369.76 & \text{meters} \end{array}$$

The second measurement (349.0 meters) has the least number of digits (one) to the right of the decimal point. So the answer must be rounded to one digit after the decimal point.

369.8 meters = 3.698×10^2 meters

Align the decimal points and subtract the numbers.

b.

$$\begin{array}{rl} 74.626 & \text{meters} \\ -\ 28.34 & \text{meters} \\ \hline 46.286 & \text{meters} \end{array}$$

The second measurement (28.34 meters) has the least number of digits (two) to the right of the decimal point. So the answer must be rounded to two digits after the decimal point.

46.29 meters = 4.629×10^1 meters

8. Perform each operation. Express your answers to the correct number of significant figures.

a. 61.2 meters + 9.35 meters + 8.6 meters

b. 9.44 meters − 2.11 meters

c. 1.36 meters + 10.17 meters

d. 34.61 meters − 17.3 meters

9. Find the total mass of three diamonds that have masses of 14.2 grams, 8.73 grams, and 0.912 gram.

Sample Problem 3.6

Significant Figures in Multiplication and Division

Perform the following operations. Give the answers to the correct number of significant figures.

a. 7.55 meters × 0.34 meter
b. 2.10 meters × 0.70 meter
c. 2.4526 meters^2 ÷ 8.4 meters
d. 0.365 meter^2 ÷ 0.0200 meter

1 Analyze **Identify the relevant concepts.** Perform the specified math operation, and then round the answer to match the measurement with the least number of significant figures.

2 Solve **Apply the concepts to this problem.**

The second measurement (0.34 meter) has the least number of significant figures (two). So the answer must be rounded to two significant figures.

a. 7.55 meters × 0.34 meter = 2.567 (meter)^2
= **2.6 meters^2**

The second measurement (0.70 meter) has the least number of significant figures (two). So the answer must be rounded to two significant figures.

b. 2.10 meters × 0.70 meter = 1.47 (meter)^2
= **1.5 meters^2**

The second measurement (8.4 meters) has the least number of significant figures (two). So the answer must be rounded to two significant figures.

c. 2.4526 meters^2 ÷ 8.4 meters = 0.291 976 meter
= **0.29 meter**

Both measurements have three significant figures. So the answer must be rounded to three significant figures.

d. 0.365 meters^2 ÷ 0.0200 meter = 18.25 meters
= **18.3 meters**

10. Solve each problem. Give your answers to the correct number of significant figures and in scientific notation.

a. 8.3 meters × 2.22 meters

b. 8432 meters^2 ÷ 12.5 meters

c. 35.2 seconds × $\frac{\text{1 minute}}{\text{60 seconds}}$

11. Calculate the volume of a warehouse that has measured dimensions of 22.4 meters by 11.3 meters by 5.2 meters. (Volume = $l \times w \times h$)

In Problem 11, the measurement with the fewest significant figures is 5.2 meters. What does this tell you?

Quick Lab

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

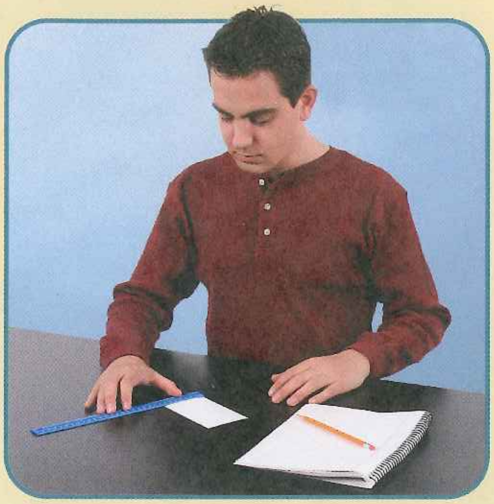
- 3-inch × 5-inch index card
- metric ruler

Accuracy and Precision

Procedure

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

2. Calculate the area ($A = l \times w$) and the perimeter [$P = 2 \times (l + w)$] of the index card. Write both your unrounded answers and your correctly rounded answers on the chalkboard.



Analyze and Conclude

1. Identify How many significant figures are in your measurements of length and of width?

2. Compare How do your measurements compare with those of your classmates?

3. Explain 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. Evaluate 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 LessonCheck

ONLINE PROBLEMS

12. Review How can you express a number in scientific notation?

13. Review How are accuracy and precision evaluated?

14. Explain Why must a given measurement always be reported to the correct number of significant figures?

15. Calculate 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.

16. Evaluate Determine the number of significant figures in each of the following measurements:

a. 11 soccer players
b. 0.070 020 meter
c. 10,800 meters
d. 0.010 square meter
e. 5.00 cubic meters
f. 507 thumbtacks

17. Calculate Solve the following and express each answer in scientific notation and to the correct number of significant figures.

a. $(5.3 \times 10^4) + (1.3 \times 10^4)$
b. $(7.2 \times 10^{-4}) \div (1.8 \times 10^3)$
c. $10^4 \times 10^{-3} \times 10^6$
d. $(9.12 \times 10^{-1}) - (4.7 \times 10^{-2})$
e. $(5.4 \times 10^4) \times (3.5 \times 10^9)$

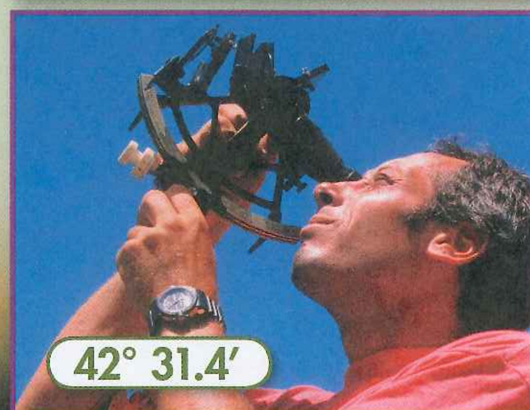
BIGIDEA QUANTIFYING MATTER

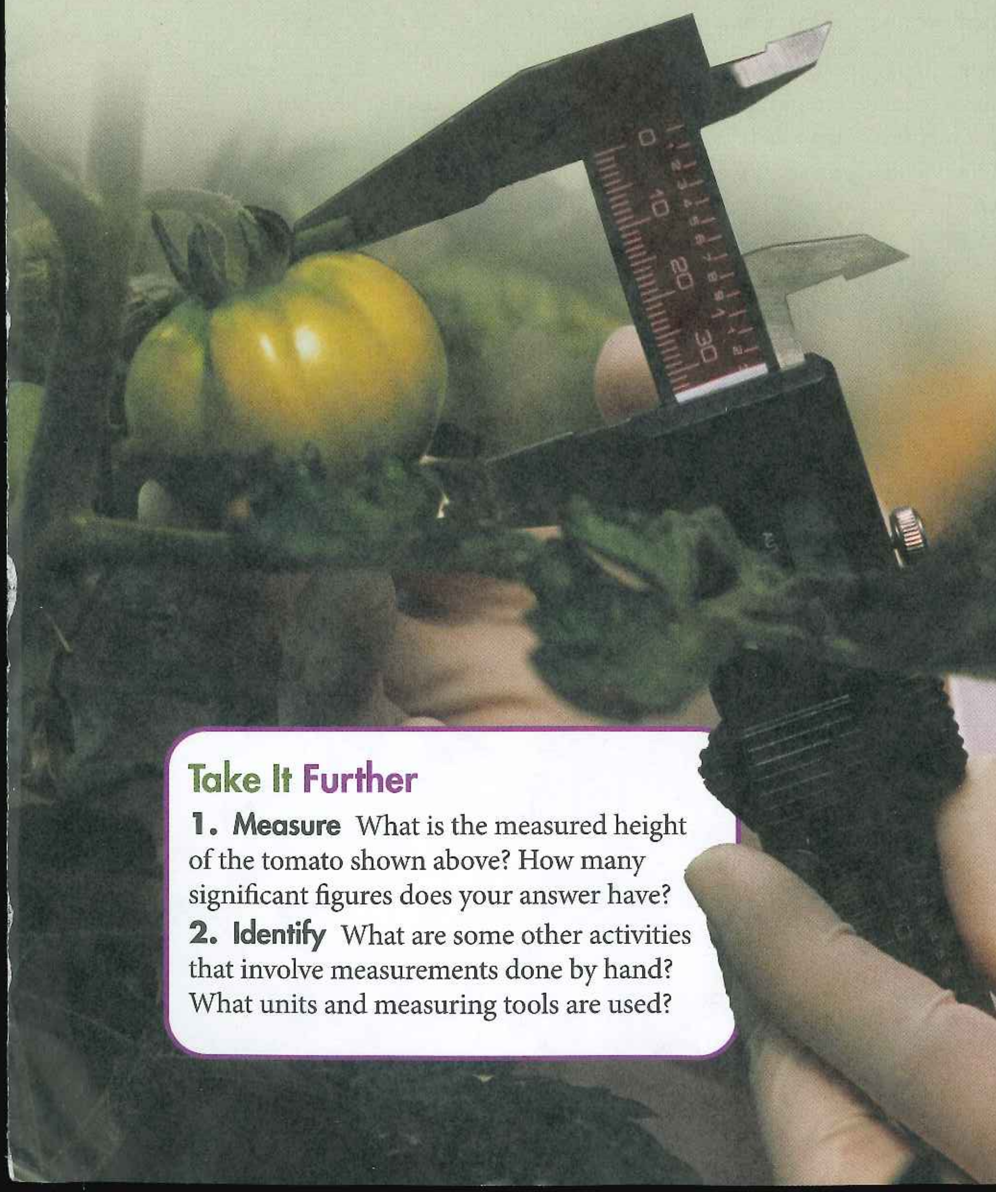
18. Write a brief paragraph explaining the differences between the accuracy, precision, and error of a measurement.

Watch What You Measure

Just because you live in a digital age doesn't mean that you no longer have to do things by hand. In fact, manually measuring quantities remains an important everyday skill in a number of professions and activities. For example, chefs measure volumes of ingredients in cups (C) or liters (L). Tailors use a tape measure calibrated in inches (in. or ″) to measure length, while biologists use metric rulers or calipers calibrated in centimeters (cm). A ship's navigator uses a sextant to measure the angle between the sun and the horizon. The angle is expressed in degrees (°) and minutes (′).

The next time you make a measurement in lab, keep in mind that lots of other measurers are rounding and noting significant figures, just like you are.





Take It Further

1. Measure What is the measured height of the tomato shown above? How many significant figures does your answer have?

2. Identify What are some other activities that involve measurements done by hand? What units and measuring tools are used?

3.2 Units of Measurement

CHEMISTRY & YOU

Q: *What's the forecast for tomorrow—hot or cold?* In the weather forecast shown here, the temperatures are in degrees, but without a temperature scale. Will the high temperature tomorrow be 28°C, which is very warm? Or 28°F, which is very cold? Without the correct units, you can't be sure. When you make a measurement, you must assign the correct units to the number. Without the units, it's impossible to communicate the measurement clearly to others.

Key Questions

What makes metric units easy to use?

What temperature units do scientists commonly use?

What determines the density of a substance?

Vocabulary

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

Learn more about **SI units** *online.*

Using SI Units

What makes metric units easy to use?

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. Derived units are used for measurements such as volume, density, and pressure.

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

SI Base Units

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

Table 3.2

Commonly Used Metric Prefixes			
Prefix	**Symbol**	**Meaning**	**Factor**
mega	M	1 million times larger than the unit it precedes	10^6
kilo	k	1000 times larger than the unit it precedes	10^3
deci	d	10 times smaller than the unit it precedes	10^{-1}
centi	c	100 times smaller than the unit it precedes	10^{-2}
milli	m	1000 times smaller than the unit it precedes	10^{-3}
micro	μ	1 million times smaller than the unit it precedes	10^{-6}
nano	n	1 billion times smaller than the unit it precedes	10^{-9}
pico	p	1 trillion times smaller than the unit it precedes	10^{-12}

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 distance race of about 42,000 m is more conveniently expressed as 42 km (42 × 1000 m). Table 3.3 summarizes the relationships among metric units of length.

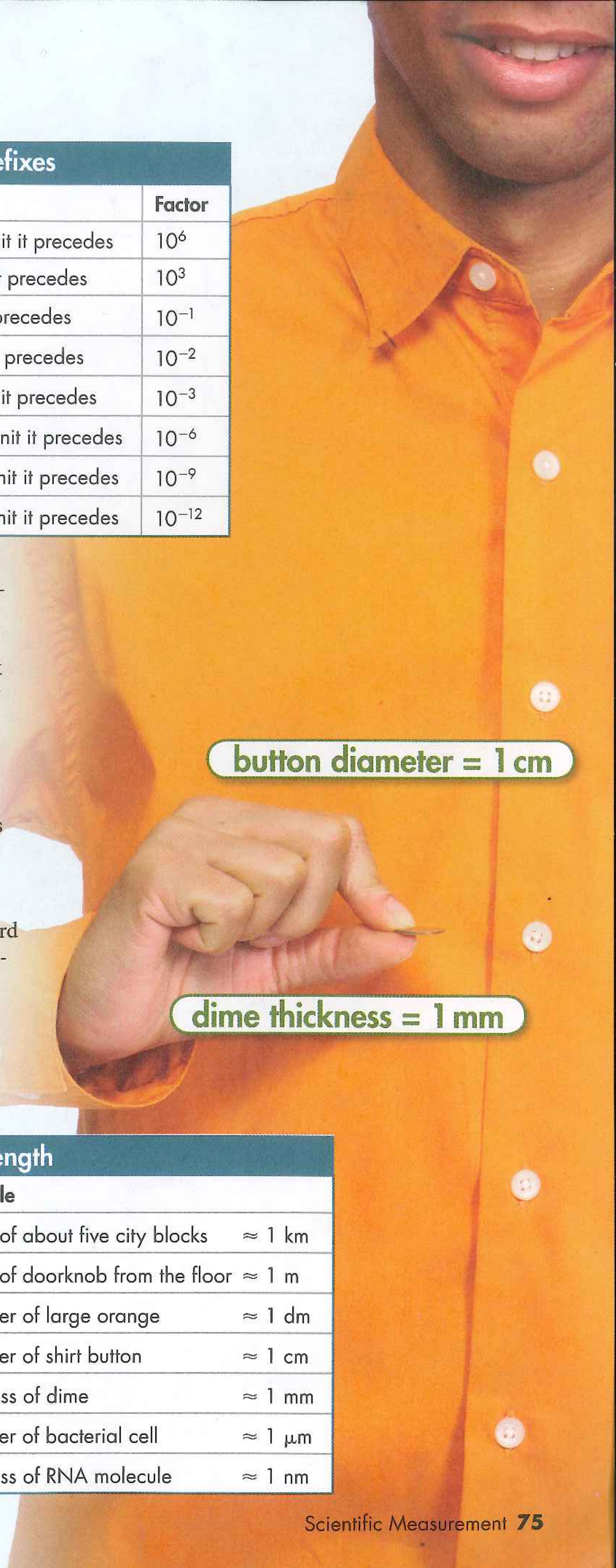
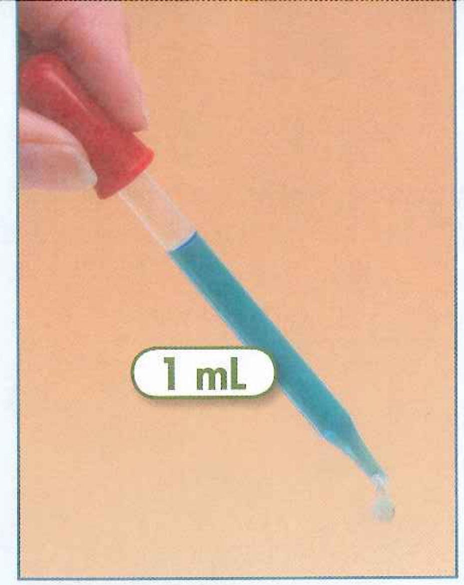


Table 3.3

Metric Units of Length			
Unit	**Symbol**	**Relationship**	**Example**
Kilometer	km	1 km = 10^3 m	length of about five city blocks ≈ 1 km
Meter	m	base unit	height of doorknob from the floor ≈ 1 m
Decimeter	dm	10^1 dm = 1 m	diameter of large orange ≈ 1 dm
Centimeter	cm	10^2 cm = 1 m	diameter of shirt button ≈ 1 cm
Millimeter	mm	10^3 mm = 1 m	thickness of dime ≈ 1 mm
Micrometer	μm	10^6 μm = 1 m	diameter of bacterial cell ≈ 1 μm
Nanometer	nm	10^9 nm = 1 m	thickness of RNA molecule ≈ 1 nm

Figure 3.5 Volumetric Units
The volume of 20 drops of liquid from a medicine dropper is about 1 mL. This is the same volume as that of a sugar cube, which is 1 cm on each edge; $1\ cm^3 = 1\ mL$. A liter bottle has a volume of 1 L, or 1000 mL.
Describe ***What is the volume of a 2-L bottle in cubic centimeters?***



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 (m^3). An automatic dishwasher has a volume of about $1\ m^3$.

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 ($10\ cm \times 10\ cm \times 10\ cm = 1000\ cm^3 = 1\ L$). A decimeter (dm) is equal to 10 cm, so 1 L is also equal to 1 cubic decimeter (dm^3). 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 $1000\ cm^3$, 1 mL and $1\ cm^3$ are the same volume. The units milliliter and cubic centimeter are thus used interchangeably. Figure 3.5 gives you some idea of the relative sizes of a liter and a milliliter. Table 3.4 summarizes the relationships among common metric 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.

Table 3.4

Metric Units of Volume

Unit	Symbol	Relationship	Example
Liter	L	base unit	quart of milk $\approx 1\ L$
Milliliter	mL	$10^3\ mL = 1\ L$	20 drops of water $\approx 1\ mL$
Cubic centimeter	cm^3	$1\ cm^3 = 1\ mL$	cube of sugar $\approx 1\ cm^3$
Microliter	μL	$10^6\ \mu L = 1\ L$	crystal of table salt $\approx 1\ \mu L$

Table 3.5

Metric Units of Mass				
Unit	**Symbol**	**Relationship**	**Example**	
Kilogram (base unit)	kg	1 kg = 10^3 g	small textbook	≈ 1 kg
Gram	g	1 g = 10^{-3} kg	dollar bill	≈ 1 g
Milligram	mg	10^3 mg = 1 g	ten grains of salt	≈ 1 mg
Microgram	μg	10^6 μg = 1 g	particle of baking powder	≈ 1 μg

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 cm^3 of water at 4°C is 1 g. 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).

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. The weight of an object can change with its location. For example, an astronaut on the surface of the moon weighs one sixth of what he weighs on Earth. The reason for this difference is that the force of Earth's gravity is about six times greater than that of the moon. The astronaut in Figure 3.6 is in free fall as he orbits Earth and is therefore weightless. Although it's possible for an object to become weightless, it can never become massless.

Figure 3.6 Weightlessness
An astronaut in orbit is weightless, but not massless. The astronaut's mass remains constant regardless of location or motion.

Units of Energy The capacity to do work or to produce heat is called **energy.** Like any other quantity, energy can be measured. The SI unit of energy is the **joule (J),** named after the English physicist James Prescott Joule (1818–1889). A common non-SI unit of energy is the calorie. 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:

$$1\ \text{J} = 0.2390\ \text{cal} \qquad 1\ \text{cal} = 4.184\ \text{J}$$

In this book, you will see energy values expressed in both joules and calories, as well as kilojoules (kJ) and kilocalories (kcal). A kilojoule is 1000 joules; a kilocalorie is 1000 calories.

Temperature Scales

What temperature units do scientists commonly use?

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. In Chapter 13, you will learn how the temperature of an object is related to the energy and motion of particles.

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 bulb 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 3.7 shows two 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.

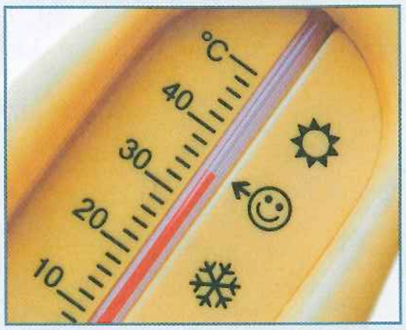


Figure 3.7 Thermometers
A bulb thermometer contains a liquid such as alcohol or mineral spirits. A dial thermometer, often used to measure the cooking temperature of meats, contains a coiled bimetallic strip.

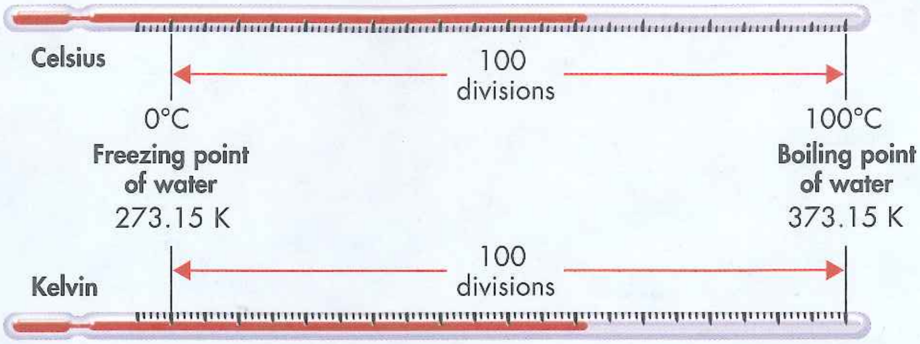


Figure 3.8 Temperature Scales
A 1°C change on the Celsius scale is equal to a 1 K change on the Kelvin scale.
Interpret Diagrams ***What is a change of 10°C equivalent to on the Kelvin scale?***

Figure 3.8 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:

$$K = °C + 273$$
$$°C = K - 273$$

CHEMISTRY & YOU

Q: *In a few countries, such as the United States, metric units are not commonly used in everyday measurements. What temperature units are used for a typical weather forecast in the United States? What about for a country that uses the metric system, such as Australia or Japan?*

Sample Problem 3.7

Converting Between Temperature Scales

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

KNOWN
Temperature in °C = 37°C

UNKNOWN
Temperature in K = ? K

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

19. The element silver melts at 960.8°C and boils at 2212°C. Express these temperatures in kelvins.

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

Figure 3.9 Floating on Water
Cranberries are less dense than water, so they float. Farmers make use of this property when it's time to harvest the crop.

Density

What determines the density of a substance?

Have you ever wondered why some objects float in water while others sink? If you think that the cranberries in Figure 3.9 float because they are lightweight, you are only partly correct. It is the relationship between the object's mass and its volume that tells you whether it will float or sink. This relationship is called density. **Density** is the ratio of the mass of an object to its volume.

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

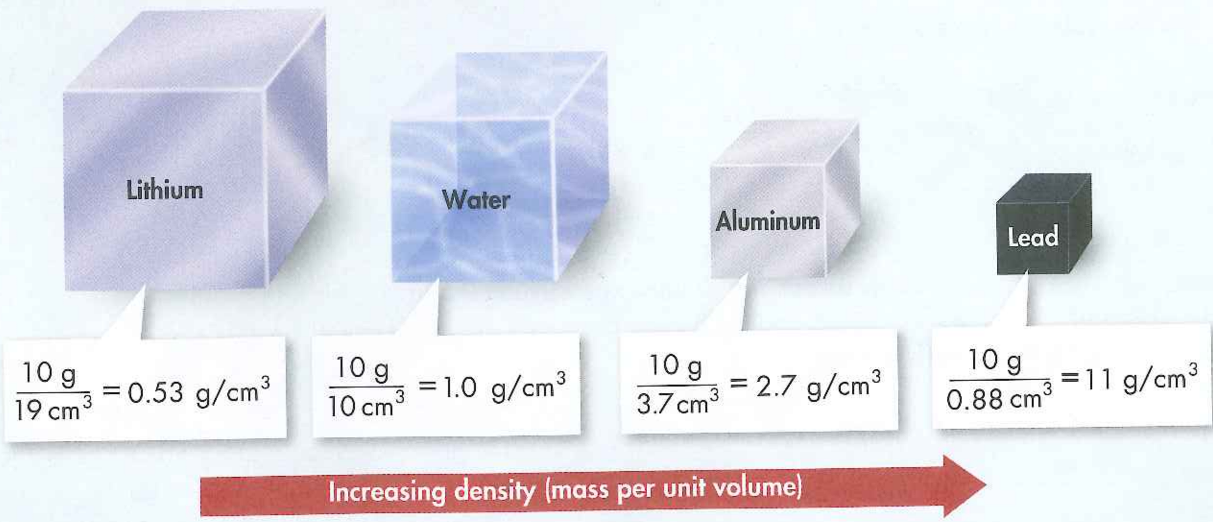
A 10.0-cm^3 piece of lead, for example, has a mass of 114 g. You can calculate the density of lead by substituting into the equation above.

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

Note that when mass is measured in grams, and volume in cubic centimeters, density has units of grams per cubic centimeter (g/cm^3). The SI unit of density is kilograms per cubic meter (kg/m^3).

Figure 3.10 compares the density of four substances: lithium, water, aluminum, and lead. 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.

Figure 3.10 Comparing Densities
A 10-g sample of pure water has less volume than 10 g of lithium, but more volume than 10 g of lead or 10 g of aluminum. The faces of the cubes are shown actual size.
Predict *Which of the solids shown will sink in water?*



Virtual Lab — Interpret Data

Densities of Some Common Materials

Solids and Liquids		Gases	
Material	**Density at 20°C (g/cm^3)**	**Material**	**Density at 20°C (g/L)**
Gold	19.3	Chlorine	2.95
Mercury	13.6	Carbon dioxide	1.83
Lead	11.3	Argon	1.66
Aluminum	2.70	Oxygen	1.33
Table sugar	1.59	Air	1.20
Corn syrup	1.35–1.38	Nitrogen	1.17
Water (4°C)	1.000	Neon	0.84
Corn oil	0.922	Ammonia	0.718
Ice (0°C)	0.917	Methane	0.665
Ethanol	0.789	Helium	0.166
Gasoline	0.66–0.69	Hydrogen	0.084

Table 3.6 Density is the mass per unit volume of a material.

a. Compare Why do you think the densities of the gases are reported in units that are different from those used for the densities of the solids and liquids?

b. Predict Would a balloon filled with carbon dioxide sink or rise in air? Explain.

c. Infer Why are the densities of corn syrup and gasoline expressed as a range of values?

Note the units here: Densities of the solids and liquids are expressed in g/cm^3. Densities of the gases are expressed in g/L.

What do you think will happen if corn oil is poured into a container of water? Using Table 3.6, you can see that the density of corn oil is less than the density of water. For that reason, the oil floats on top of the water. Figure 3.11 shows different liquids forming distinct layers in a container due to differences in density. For example, the corn syrup (colored red), sinks below the water (colored green) because the density of corn syrup is greater than the density of water.

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.

What happens to the density of a substance as its temperature increases? 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 (0.917 g/cm^3) than liquid water (1.000 g/cm^3).

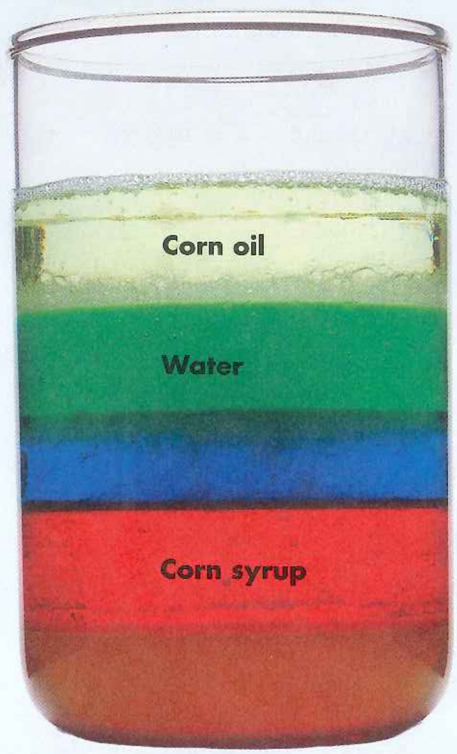


Figure 3.11 Liquid Layers
Because of differences in density, the liquids separate into layers.
Compare *Is the blue-colored liquid more or less dense than water?*

Sample Problem 3.8

Calculating Density

A copper penny has a mass of 3.1 g and a volume of 0.35 cm^3. What is the density of copper?

1 Analyze **List the knowns and the unknown.** Use the known values and the equation for density to solve the problem.

KNOWNS	UNKNOWN
mass = 3.1 g	density = ? g/cm^3
volume = 0.35 cm^3	

2 Calculate **Solve for the unknown.**

Start with the equation for density.

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

The calculated answer must be rounded to two significant figures.

Substitute the known values for mass and volume and then calculate.

$$\text{Density} = \frac{3.1\ \text{g}}{0.35\ \text{cm}^3} = 8.8571\ \text{g/cm}^3 = 8.9\ \text{g/cm}^3$$

3 Evaluate **Does the result make sense?** A piece of copper with a volume of about 0.3 cm^3 has a mass of about 3 grams. About three times that volume of copper, 1 cm^3, should have a mass three times larger, about 9 grams. This estimate is close to the calculated result.

21. 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 cm^3 and a mass of 612 g. Calculate the density. Is the metal aluminum?

22. A bar of silver has a mass of 68.0 g and a volume of 6.48 cm^3. What is the density of silver?

ONLINE PROBLEMS

3.2 LessonCheck

23. **Review** Why are metric units easy to use?

24. **Identify** What temperature units do scientists commonly use?

25. **Review** What determines density?

26. **Identify** Write the name and symbol of the SI units for mass, length, volume, and temperature.

27. **Define** Write the symbol and meaning of each prefix below.

a. *milli-* **b.** *nano-* **c.** *deci-* **d.** *centi-*

28. **List** Arrange the following units in order from largest to smallest: m^3, mL, cL, μL, L, dL.

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

30. **Compare** State the difference between mass and weight.

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

32. **Calculate** A weather balloon is inflated to a volume of 2.2×10^3 L with 374 g of helium. What is the density of helium in grams per liter?

33. **Apply Concepts** A 68-g bar of gold is cut into three equal pieces. How does the density of each piece compare to the density of the original gold bar?

34. **Interpret Data** Look up the densities of the elements in Group 1A on page R2. Which Group 1A elements are less dense than pure water at 4°C?

35. **Explain** How does density vary with temperature?

Carbon Footprints

To measure a footprint, you might use units such as centimeters or inches. But what about a carbon footprint? A carbon footprint is a measure of how much greenhouse gas is released into the atmosphere by a person, country, or industry. Greenhouse gases, such as carbon dioxide (CO_2) and methane (CH_4), are gases that contribute to global warming.

Any activity that involves the burning of fossil fuels results in carbon dioxide emissions. Car travel, air travel, home heating and cooling, and electricity usage all add to an individual's carbon footprint. Your own carbon footprint is the total mass of CO_2 that you put into the atmosphere over the course of a year. This quantity can be expressed in metric tons (t) of CO_2 per year. A metric ton equals 1000 kg. So, the units of your carbon footprint can be abbreviated as: t CO_2/yr or 10^3 kg CO_2/yr.

CARBON COSTS Your choices affect the size of your carbon footprint. For example, using a clothes dryer consumes electricity, but hanging wet laundry on a clothesline does not. The things you buy also contribute to your carbon footprint. Not only is energy required to make these goods, but the goods themselves (such as TVs) may consume energy.

FOOTPRINT UNITS The carbon footprint of fresh produce can be expressed in g CO_2 per serving. Cars require different units: kg CO_2 per gallon of gasoline. For planes, the units are kg CO_2 per passenger mile.

Take It Further

1. Calculate A car emits 8.6 kg of CO_2 per gallon of unleaded gas. How much CO_2 is produced if the car burns 2.5 gal of fuel?

2. Infer What factors do you think determine the carbon footprint of an apple? Why might the carbon footprints of two apples in the same store differ substantially?

3.3 Solving Conversion Problems

CHEMISTRY & YOU

Q: *How can you convert U.S. dollars to euros?* Perhaps you have traveled to another country 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 an important part of having a good trip. After all, you often must use cash to pay for meals, transportation, and souvenirs. Because each country's currency compares differently with the U.S. dollar, knowing how to convert currency units correctly is essential. Conversion problems are readily solved by a problem-solving approach called dimensional analysis.

Key Questions

What happens when a measurement is multiplied by a conversion factor?

What kinds of problems can you solve using dimensional analysis?

Vocabulary

- conversion factor
- dimensional analysis

Conversion Factors

What happens when a measurement is multiplied by a conversion factor?

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 meaurements 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.

$$\frac{1\text{m}}{1\text{m}} = \frac{100\text{ cm}}{1\text{ m}} = 1 \qquad \text{or} \qquad \frac{1\text{ m}}{100\text{ cm}} = \frac{100\text{ cm}}{100\text{ cm}} = 1$$

conversion factors

The ratios 100 cm/1 m and 1 m/100 cm are examples of conversion factors. A **conversion factor** is a ratio of equivalent measurements. The measurement in the numerator (on the top) is equivalent to the measurement in the denominator (on the bottom). The conversion factors shown above are read "one hundred centimeters per meter" and "one meter per hundred centimeters."

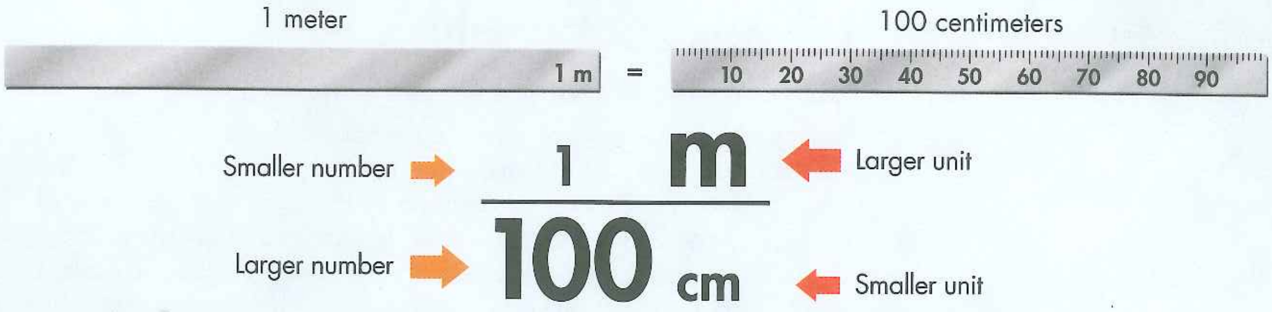


Figure 3.12 Conversion Factor
The two parts of a conversion factor, the numerator and the denominator, are equal.

See conversion factors animated online. KINETIC ART

Figure 3.12 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

$$\frac{1000\text{ g}}{1\text{ kg}} \quad \text{and} \quad \frac{1\text{ kg}}{1000\text{ g}}$$

Figure 3.13 shows a scale that can be used to measure mass in grams or kilograms. If you read the scale in terms of grams, you can convert the mass to kilograms by multiplying by the conversion factor 1 kg/1000 g.

The relationship between nanometers and meters is given by the equation 10^9 nm = 1 m. The possible conversion factors are

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

Common volumetric units used in chemistry include the liter and the microliter. The relationship 1 L = 10^6 μL 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 have learned about metric prefixes, you should easily be able to write conversion factors that relate equivalent metric quantities.

Figure 3.13 Measuring Mass
This scale is calibrated to measure mass to the nearest 20 g.
Interpret Photos *What is the scale showing in grams? In kilograms?*

Dimensional Analysis

What kinds of problems can you solve using dimensional analysis?

Some problems are best solved using algebra. For example, converting a kelvin temperature to Celsius can be done by using the equation °C = K − 273. Many problems in chemistry are conveniently solved using 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 technique is to use it to solve an everyday situation, as in Sample Problem 3.9.

As you read Sample Problem 3.10, you might see how the same problem could be solved algebraically but is more easily solved using dimensional analysis. In either case, you should choose the problem-solving method that works best for you. Try to be flexible in your approach to problem solving, as no single method is best for solving every type of problem.

Sample Problem 3.9

Using Dimensional Analysis

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

1 Analyze List the knowns and the unknown.
To convert time in hours to time in seconds, you'll need two conversion factors. First you must convert hours to minutes: h ⟶ min. Then you must convert minutes to seconds: min ⟶ s. Identify the proper conversion factors based on the relationships 1 h = 60 min and 1 min = 60 s.

KNOWNS	UNKNOWN
time worked = 8 h	seconds worked = ? s
1 hour = 60 min	
1 minute = 60 s	

2 Calculate Solve for the unknown.

The first conversion factor is based on 1 h = 60 min. The unit hours must be in the denominator so that the known unit will cancel.

$$\frac{60 \text{ min}}{1 \text{ h}}$$

The second conversion factor is based on 1 min = 60 s. The unit minutes must be in the denominator so that the desired units (seconds) will be in your answer.

$$\frac{60 \text{ s}}{1 \text{ min}}$$

Multiply the time worked by the conversion factors.

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

Before you do the actual arithmetic, it's 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.

3 Evaluate Does the result make sense? The answer has the desired unit (s). Since the second is a small unit of time, you should expect a large number of seconds in 8 hours. The answer is exact since the given measurement and each of the conversion factors is exact.

36. How many minutes are there in exactly one week?

37. How many seconds are in an exactly 40-hour work week?

Sample Problem 3.10

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.
From the known mass of copper, use the appropriate conversion factor to calculate the number of students who can do the experiment. The desired conversion is mass of copper ⟶ number of students.

KNOWNS
mass of copper available = 50.0 g Cu
Each student needs 1.84 grams of copper.

UNKNOWN
number of students = ?

2 Calculate Solve for the unknown.

The experiment calls for 1.84 grams of copper per student. Based on this relationship, you can write two conversion factors.

$$\frac{1.84 \text{ g Cu}}{1 \text{ student}} \quad \text{and} \quad \frac{1 \text{ student}}{1.84 \text{ g Cu}}$$

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

Because the desired unit for the answer is students, use the second conversion factor. Multiply the mass of copper by the conversion factor.

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

3 Evaluate Does the result make sense? The unit of the answer (students) is the one desired. 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.

38. 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?

Here's a tip: The equalities needed to write a particular conversion factor may be given in the problem. In other cases, you'll need to know or look up the necessary equalities.

39. 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 in °F?

40. 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?

CHEMISTRY & YOU

Q: *Look up the exchange rate between U.S. dollars and euros on the Internet. Write a conversion factor that allows you to convert from U.S. dollars to euros. How many euros could you buy with $50?*

Simple Unit Conversions In chemistry, as in everyday life, you often need to express a measurement in a unit different from the one given or measured initially. **Dimensional analysis is a powerful tool for solving conversion problems in which a measurement with one unit is changed to an equivalent measurement with another unit.** Sample Problems 3.11 and 3.12 walk you through how to solve simple conversion problems using dimensional analysis.

Sample Problem 3.11

Converting Between Metric Units

Express 750 dg in grams. (Refer to Table 3.2 if you need to refresh your memory of metric prefixes.)

1 Analyze **List the knowns and the unknown.**
The desired conversion is decigrams ⟶ grams. Multiply the given mass by the proper conversion factor.

KNOWNS	UNKNOWN
mass = 750 dg	mass = ? g
1 g = 10 dg	

2 Calculate **Solve for the unknown.**

Use the relationship 1 g = 10 dg to write the correct conversion factor. → $\frac{1\ \text{g}}{10\ \text{dg}}$

Multiply the known mass by the conversion factor. → $750\ \cancel{\text{dg}} \times \frac{1\ \text{g}}{10\ \cancel{\text{dg}}} = 75\ \text{g}$

Note that the known unit (dg) is in the denominator and the unknown unit (g) is in the numerator.

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 answer has the correct unit (g) and the correct number of significant figures.

41. 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

42. Convert the following:
- **a.** 15 cm^3 to liters
- **b.** 7.38 g to kilograms
- **c.** 6.7 s to milliseconds
- **d.** 94.5 g to micrograms

CHEM TUTOR

Sample Problem 3.12

Using Density as a Conversion Factor

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/cm^3.

1 Analyze List the knowns and the unknown. You need to convert the mass of the coin into a corresponding volume. The density gives you the following relationship between volume and mass: 1 cm^3 Ag = 10.5 g Ag. Multiply the given mass by the proper conversion factor to yield an answer in cm^3.

KNOWNS	UNKNOWN
mass = 14 g	volume of coin = ? cm^3
density of silver = 10.5 g/cm^3	

2 Calculate Solve for the unknown.

Use the relationship 1 cm^3 Ag = 10.5 g Ag to write the correct conversion factor.

$$\frac{1\ cm^3\ Ag}{10.5\ g\ Ag}$$

Notice that the known unit (g) is in the denominator and the unknown unit (cm^3) is in the numerator.

Multiply the mass of the coin by the conversion factor.

$$14\ \cancel{g\ Ag} \times \frac{1\ cm^3\ Ag}{10.5\ \cancel{g\ Ag}} = 1.3\ cm^3\ Ag$$

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

43. Use dimensional analysis and the given densities to make the following conversions:

a. 14.8 g of boron to cm^3 of boron. The density of boron is 2.34 g/cm^3.

b. 4.62 g of mercury to cm^3 of mercury. The density of mercury is 13.5 g/cm^3.

44. Rework the preceding problems by applying the following equation:

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

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

Multistep Problems Many complex tasks in your life are best handled by breaking them down into smaller, 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 Problems 3.13 and 3.14 illustrate the use of multiple conversion factors.

Sample Problem 3.13

Converting Between Metric Units

The diameter of a sewing needle is 0.073 cm. What is the diameter in micrometers?

1 Analyze List the knowns and the unknown. The desired conversion is centimeters ⟶ micrometers. The problem can be solved in a two-step conversion. First change centimeters to meters; then change meters to micrometers: centimeters ⟶ meters ⟶ micrometers.

KNOWNS	UNKNOWN
diameter = 0.073 cm = 7.3×10^{-2} cm	diameter = ? μm
10^2 cm = 1 m	
1 m = 10^6 μm	

2 Calculate Solve for the unknown.

Use the relationship 10^2 cm = 1 m to write the first conversion factor.

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

Use the relationship 1 m = 10^6 μm to write the second conversion factor.

$$\frac{10^6\ \mu\text{m}}{1\ \text{m}}$$

Each conversion factor is written so that the unit in the denominator cancels the unit in the numerator of the previous factor.

Multiply the known length by the conversion factors.

$$7.3 \times 10^{-2}\ \cancel{\text{cm}} \times \frac{1\ \cancel{\text{m}}}{10^2\ \cancel{\text{cm}}} \times \frac{10^6\ \mu\text{m}}{1\ \cancel{\text{m}}} = 7.3 \times 10^2\ \mu\text{m}$$

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.

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

47. The diameter of Earth is 1.3×10^4 km. What is the diameter expressed in decimeters?

Sample Problem 3.14

Converting Ratios of Units

The density of manganese, a metal, is 7.21 g/cm^3. What is the density of manganese expressed in units of kg/m^3?

1 Analyze List the knowns and the unknown.

The desired conversion is $g/cm^3 \longrightarrow kg/m^3$. The mass unit in the numerator must be changed from grams to kilograms: $g \longrightarrow kg$. In the denominator, the volume unit must be changed from cubic centimeters to cubic meters: $cm^3 \longrightarrow m^3$. Note that the relationship $10^6\ cm^3 = 1\ m^3$ was derived by cubing the relationship $10^2\ cm = 1\ m$. That is, $(10^2\ cm)^3 = (1\ m)^3$, or $10^6\ cm^3 = 1\ m^3$.

KNOWNS

density of manganese = 7.21 g/cm^3

$10^3\ g = 1\ kg$

$10^6\ cm^3 = 1\ m^3$

UNKNOWN

density of manganese = ? kg/m^3

2 Calculate Solve for the unknown.

Multiply the known density by the correct conversion factors.

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

3 Evaluate Does the result make sense?

Because the physical size of the volume unit m^3 is so much larger than 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.

48. Gold has a density of 19.3 g/cm^3. What is the density in kilograms per cubic meter?

49. There are 7.0×10^6 red blood cells (RBCs) in 1.0 mm^3 of blood. How many red blood cells are in 1.0 L of blood?

3.3 LessonCheck

50. Review 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?

51. Review What types of problems can be solved using dimensional analysis?

52. Identify 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

53. Calculate Make the following conversions. Express your answers in scientific notation.

a. 14.8 g = ? μg

b. 3.72 g = ? kg

c. 66.3 L = ? cm^3

d. 7.5×10^4 J = ? kJ

e. 3.9×10^5 mg = ? dg

f. 2.1×10^{-4} dL = ? μL

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

55. Apply Concepts Light travels at a speed of 3.00×10^{10} cm/s. What is the speed of light in kilometers/hour?

Small-Scale Lab

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-1982 penny
- a 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.

Analyze

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

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

2. Calculate The density of water is 1.00 g/cm^3. Calculate the volume of a single drop in cm^3 and mL. (1 mL = 1 cm^3) What is the volume of a drop in microliters (μL)? (1000 μL = 1 mL)

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

4. Calculate 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. Calculate 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. Explain 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 an Experiment 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 Data 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 an Experiment 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 an Experiment Design and carry out some experiments to determine the volume of liquid that an aluminum soda can will hold.

5. Design an Experiment 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 an Experiment 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?

3 Study Guide

BIGIDEA QUANTIFYING MATTER

Scientists express the degree of uncertainty in their measurements and calculations by using significant figures. In general, a calculated answer cannot be more precise than the least precise measurement from which it was calculated. Dimensional analysis is a problem-solving method that involves analyzing the units of the given measurement and the unknown to plan a solution.

3.1 Using and Expressing Measurements

In scientific notation, the coefficient is always a number greater than or equal to one and less than ten. The exponent is an integer.

To evaluate accuracy, 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.

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.

- measurement (62)
- scientific notation (62)
- accuracy (64)
- precision (64)
- accepted value (65)
- experimental value (65)
- error (65)
- percent error (65)
- significant figures (66)

Key Equations

$$\text{Error} = \text{experimental value} - \text{accepted value}$$

$$\text{Percent error} = \frac{|\text{error}|}{\text{accepted value}} \times 100\%$$

3.2 Units of Measurement

All metric units are based on multiples of 10. As a result, you can convert between units easily.

Scientists commonly use two equivalent units of temperature, the degree Celsius and the kelvin.

Density is an intensive property that depends only on the composition of a substance, not on the size of the sample.

- International System of Units (SI) (74)
- meter (m) (75)
- liter (L) (76)
- kilogram (kg) (77)
- gram (g) (77)
- weight (77)
- energy (77)
- joule (J) (77)
- calorie (cal) (77)
- temperature (78)
- Celsius scale (78)
- Kelvin scale (78)
- absolute zero (79)
- density (80)

Key Equations

$$K = {}^\circ C + 273$$

$${}^\circ C = K - 273$$

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

3.3 Solving Conversion Problems

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.

Dimensional analysis is a powerful tool for solving conversion problems in which a measurement with one unit is changed to an equivalent measurement with another unit.

- conversion factor (84)
- dimensional analysis (86)

Math Tune-Up: Conversion Problems

Problem	❶ Analyze	❷ Calculate	❸ Evaluate
A grocer is selling oranges at 3 for $2. How much would it cost to buy a dozen oranges?	Knowns: 3 oranges = $2 1 dozen = 12 Unknown: Cost of 12 oranges = ? The desired conversion is oranges ⟶ $.	Use the relationship 3 oranges = $2 to write the correct conversion factor. $\frac{\$2}{3 \text{ oranges}}$ Multiply the known quantity by the conversion factor. $12 \cancel{\text{oranges}} \times \frac{\$2}{3 \cancel{\text{oranges}}} = \$8$	A dozen is larger than the number 3, so the cost should exceed $2. The known unit (oranges) cancels, and the answer has the correct unit ($).
Convert the volume 865 cm^3 to liters.	Knowns: volume = 865 cm^3 10^3 cm^3 = 1 L Unknown: volume = ? L The desired conversion is cm^3 ⟶ L.	Use the relationship 10^3 cm^3 = 1 L to write the correct conversion factor. $\frac{1 \text{ L}}{10^3 \text{ cm}^3}$ Multiply the known volume by the conversion factor. $865 \cancel{\text{cm}^3} \times \frac{1 \text{ L}}{10^3 \cancel{\text{cm}^3}} = 0.865 \text{ L}$	A cubic centimeter is much smaller than a liter, so the answer should be numerically smaller than the given measurement. The known unit (cm^3) cancels, and the answer has the correct unit (L).
Express the length 8.2×10^{-4} μm in centimeters.	Knowns: length = 8.2×10^{-4} μm 10^6 μm = 1 m 1 m = 10^2 cm Unknown: length = ? L The desired conversion is μm ⟶ cm. First change μm to m; then change m to cm: μm ⟶ m ⟶ cm	Use the relationship 10^6 μm = 1 m to write the first conversion factor. $\frac{1 \text{ m}}{10^6 \text{ μm}}$ Use the relationship 1 m = 10^2 cm to write the second conversion factor. $\frac{10^2 \text{ cm}}{1 \text{ m}}$ Multiply the known length by the conversion factors. $8.2 \times 10^{-4} \cancel{\text{μm}} \times \frac{1 \cancel{\text{m}}}{10^6 \cancel{\text{μm}}} \times \frac{10^2 \text{ cm}}{1 \cancel{\text{m}}}$ $= 8.2 \times 10^{-8} \text{ cm}$	A micrometer is smaller than a centimeter, so the answer should be numerically smaller than the given measurement. The known unit (μm) cancels, and the answer has the correct unit (cm). **Hint:** For a multistep problem, take it one conversion at a time.

3 Assessment

★ Solutions appear in Appendix E

Lesson by Lesson

3.1 Using and Expressing Measurements

56. 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.

Mass of Cylinder (g)			
	Colin	**Lamont**	**Kivrin**
Weighing 1	47.13	47.45	47.95
Weighing 2	47.94	47.39	47.91
Weighing 3	46.83	47.42	47.89
Weighing 4	47.47	47.41	47.93

57. How many significant figures are in each underlined measurement?

a. <u>60 s</u> = 1 min
b. <u>47.70 g</u> of copper
c. 1 km = <u>1000 m</u>
d. <u>25 computers</u>
e. <u>9 innings</u> in a baseball game
f. <u>0.0950 m</u> of gold chain

58. 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
e. 0.007 498 3 $\times$ 10^4 mm
f. 1764.9 mL

★**59.** Round off each of the answers correctly.

a. 8.7 g + 15.43 g + 19 g = 43.13 g
b. 4.32 cm $\times$ 1.7 cm = 7.344 cm^2
c. 853.2 L − 627.443 L = 225.757 L
d. 38.742 m^2 ÷ 0.421 m = 92.023 75 m
e. 5.40 m $\times$ 3.21 m $\times$ 1.871 m = 32.431 914 m^3
f. 5.47 m^3 + 11 m^3 + 87.300 m^3 = 103.770 m^3

★**60.** Express each of the rounded-off answers in Problems 58 and 59 in scientific notation.

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

3.2 Units of Measurement

62. Write the SI base unit of measurement for each of these quantities.

a. time
b. length
c. temperature
d. mass
e. energy
f. amount of substance

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

64. 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

65. State the relationship between degrees Celsius and kelvins.

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

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

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

★**69.** A shiny, gold-colored bar of metal weighing 57.3 g has a volume of 4.7 cm^3. Is the bar of metal pure gold?

70. Three balloons filled with neon, carbon dioxide, and hydrogen are released into the atmosphere. Using the data in Table 3.6 on page 81, describe the movement of each balloon.

3.3 Solving Conversion Problems

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

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

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

*74. 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. 8.25×10^2 cg to nanograms

*75. Make the following conversions:

a. 0.44 mL/min to microliters per second
b. 7.86 g/cm^2 to milligrams per square millimeter
c. 1.54 kg/L to grams per cubic centimeter

76. How many milliliters are contained in 1 m^3?

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

mg	g	cg	kg
a. ____	b. ____	28.3	c. ____
6.6×10^3	d. ____	e. ____	f. ____
g. ____	2.8×10^{-4}	h. ____	i. ____

Understand Concepts

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

79. Rank these numbers from smallest to largest.

a. 5.3×10^4
b. 57×10^3
c. 4.9×10^{-2}
d. 0.0057
e. 5.1×10^{-3}
f. 0.0072×10^2

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.

Example	°C	°F
Melting point of selenium	221	430
Boiling point of water	100	212
Normal body temperature	37	98.6
Freezing point of water	0	32
Boiling point of chlorine	−34.6	−30.2

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

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

84. 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?

85. Write six conversion factors involving these units of measure: $1 \text{ g} = 10^2 \text{ cg} = 10^3 \text{ mg}$.

*86. 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.

*87. The density of dry air measured at 25°C is 1.19×10^{-3} g/cm^3. What is the volume of 50.0 g of air?

88. 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/cm^3.

*89. 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?

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

*91. 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 1.38×10^4 g. What is the density of olive oil in kilograms per liter?

92. 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 the number of carbon atoms on the x-axis.

Boiling point (°C)	Number of carbon atoms
−162.0	1
−42.0	3
36.0	5
98.0	7
151.0	9

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?

***93.** Earth is approximately 1.5×10^8 km from the sun. How many minutes does it take light to travel from the sun to Earth? The speed of light is 3.0×10^8 m/s.

***94.** The average density of Earth is 5.52 g/cm^3. Express this density in units of kg/dm^3.

95. 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?

Think Critically

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

***97. Calculate** 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?

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

99. Evaluate 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.

100. Describe 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 think of situations of more consequence?

101. Graph 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.

Volume of sulfur (cm^3)	Mass of sulfur (g)
11.4	23.5
29.2	60.8
55.5	115
81.1	168

102. Analyze Data 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.

Enrichment

*103. **Calculate** The mass of a cube of iron is 355 g. Iron has a density of 7.87 g/cm^3. What is the mass of a cube of lead that has the same dimensions?

*104. **Calculate** Sea water contains 8.0×10^{-1} 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. **Calculate** 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. **Graph** 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.

Volume (mL)	Mass (g)
14	103.0
27	120.4
41	139.1
55	157.9
82	194.1

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

*107. **Predict** A 34.5-g gold nugget is dropped into a graduated cylinder containing water. By how many milliliters does the measured volume increase if the nugget is completely covered by water? The density of water is 1.0 g/mL. The density of gold is 19.3 g/cm^3.

108. **Predict** Equal amounts of mercury, water, and corn oil are added to a beaker. Use Table 3.6 to help you anwer the following questions.

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?

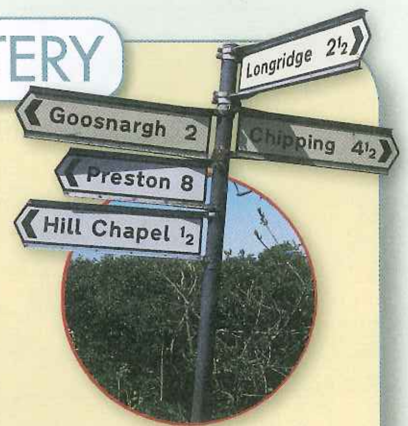
Write About Science

109. **Describe** For one of the topics below, write a short paragraph that identifies both metric and nonmetric units that are commonly used to communicate information.

a. measurements used in cooking
b. measurements used in sports
c. measurements used in transportation

110. **Connect to the BIGIDEA** Explain how the three-step problem-solving approach defined in Chapter 1 (*Analyze, Calculate, Evaluate*) applies to problems that involve dimensional analysis.

CHEMYSTERY



Just Give Me a Sign

The road signs point to locations in England. Although England has adopted metric units for many everyday quantities, distances shown on road signs are not among them. The road signs above list distances in miles, a nonmetric unit. Speed-limit signs in England are typically expressed in miles per hour—also nonmetric. However, in the same country, gasoline is sold by metric units of volume (liters), fabric is measured in metric units of area (square meters), and the local weather report uses metric units of temperature (°C).

111. **Calculate** The relationship between kilometers and miles (mi) is 1 km = 0.621 mi. How far is it to Chipping in kilometers?

112. **Calculate** Suppose you encounter the road signs above while cycling. If your average speed is 18 km/h, how many minutes will it take you to reach Preston?

113. **Connect to the BIGIDEA** Describe two ways in which the road signs above might be considered examples of "uncertainty in measurement."

Standardized Test Prep

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, cL

2. Which answer represents the measurement 0.00428 g rounded to two significant figures?
   (A) 4.28×10^3 g
   (B) 4.3×10^3 g
   (C) 4.3×10^{-3} g
   (D) 4.0×10^{-3} 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 $\mu m^3 = 1\ cm^3$?
   (A) 10^4
   (B) 10^6
   (C) 10^8
   (D) 10^{12}

5. 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 6–7, identify the known and the unknown. Include units in your answers.

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

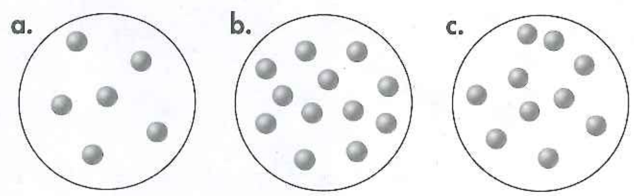
7. A graduated cylinder contains 44.2 mL of water. A 48.6-g piece of metal is carefully dropped into the cylinder. When the metal is completely covered with water, the water rises to the 51.3-mL mark. What is the density of the metal?

Tips for Success

Interpret Diagrams Before you answer questions about a diagram, study the diagram carefully. Ask: What is the diagram showing? What does it tell me?

Use the diagrams below to answer Questions 8 and 9.

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.



8. List the windows in order of decreasing density.

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

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		Statement II
10.	There are five significant figures in the measurement 0.00450 m.	BECAUSE	All zeros to the right of a decimal point in a measurement are significant.
11.	Precise measurements will always be accurate measurements.	BECAUSE	A value that is measured 10 times in a row must be accurate.
12.	A temperature in kelvins is always numerically larger than the same temperature in degrees Celsius.	BECAUSE	A temperature in kelvins equals a temperature in degrees Celsius plus 273.

If You Have Trouble With . . .

Question	1	2	3	4	5	6	7	8	9	10	11	12
See Lesson	3.2	3.1	3.3	3.3	3.2	3.3	3.2	3.2	3.2	3.1	3.1	3.2