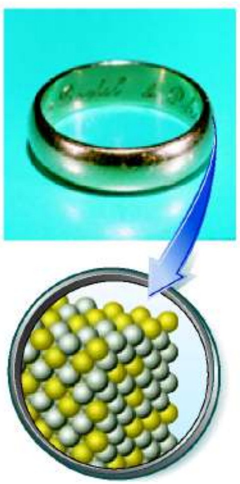
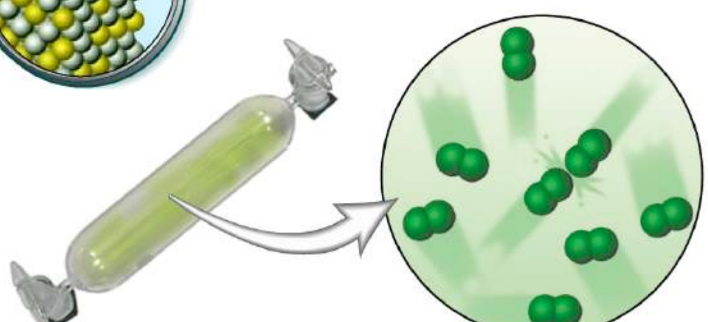
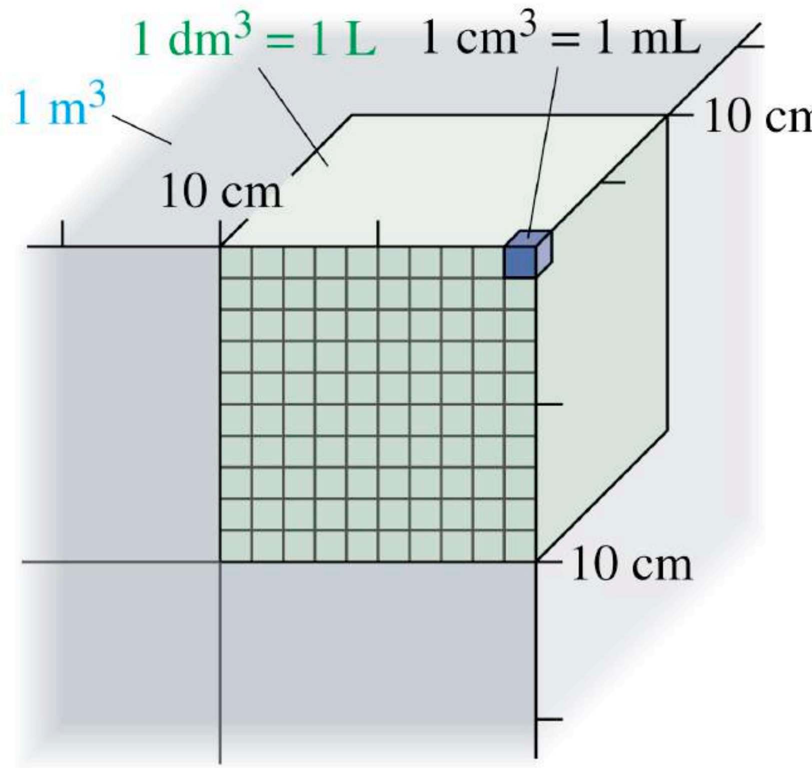
Chapter One

Chemistry: Matter and Measurement







Getting Started: Some Key Terms

- **Chemistry** is the study of the composition, structure, and properties of matter and of changes that occur in matter.
- **Matter** is anything that has mass and occupies space.
  * Matter is the stuff that things are made of.

Key Terms

- **Atoms** are the smallest distinctive units in a sample of matter.
- **Molecules** are larger units in which two or more atoms are joined together.
  - * Examples: Water consists of molecules, each having two atoms of hydrogen and one of oxygen.
  - * Oxygen *gas* consists of molecules, each having two atoms of oxygen.

Water molecule

Oxygen molecule

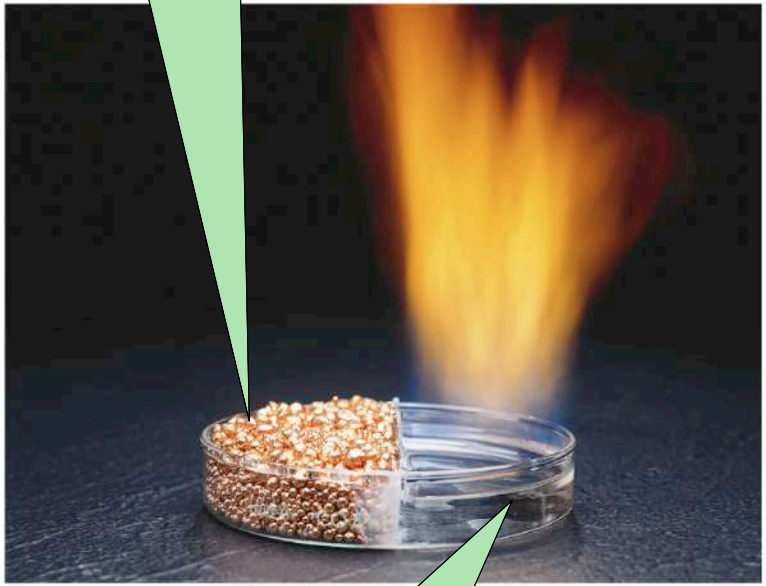
Key Terms

- **Composition** – the types of atoms and their relative proportions in a sample of matter.
- The composition of water is two parts (by atoms) of hydrogen to one part (by atoms) of oxygen.
- The composition of water is 11.2% hydrogen by *mass*, 88.8% oxygen by *mass*.
  - (Why the difference? Because hydrogen atoms and oxygen atoms don't have the same mass!)
  - More on mass composition in Chapter 3.

Key Terms: Properties

- A **physical property** is displayed by a sample of matter without undergoing any change in the composition of the matter.
  - * Physical properties include mass, color, volume, temperature, density, melting point, etc.
- **Chemical property** – displayed by a sample of matter as it undergoes a change in composition.
  - * Flammability, toxicity, reactivity, acidity are all chemical properties.

Copper is red-brown, opaque, solid: *physical* properties.



Ethanol is flammable: a *chemical* property.

Key Terms: Properties

- In a **physical change**, there is no change in composition.
- No new substances are formed.
  - Examples include: evaporation; melting; cutting a piece of wood; dissolving sugar in water.
- In a **chemical change** or *chemical reaction*, the matter undergoes a change in composition.
- New substances are formed.
  - Examples include: burning gasoline; dissolving metal in acid; spoilage of food.

The vapor burns, combining with oxygen: a *chemical* change.

The liquid fuel evaporates: a *physical* change.

Classifying Matter

Figure 1.3

Atoms — Molecules

make up

ALL MATTER

which exists as

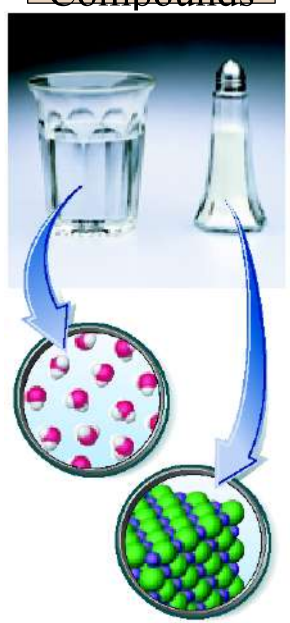
Substances — Mixtures

Substances, which may be: Elements, Compounds

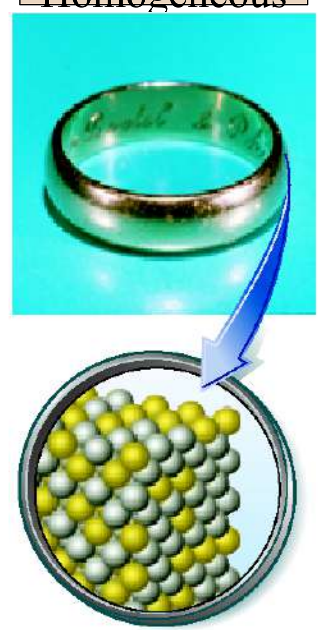
Mixtures, which may be: Homogeneous, Heterogeneous

Elements

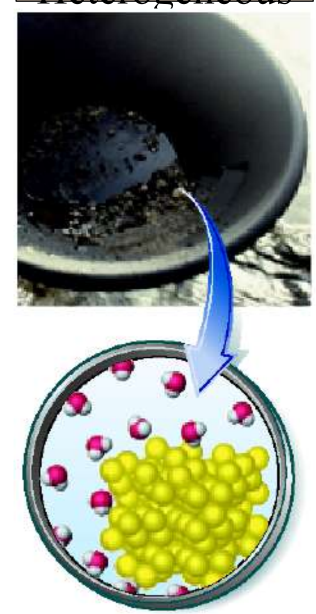
Compounds



Homogeneous



Heterogeneous



Classifying Matter

- A **substance** has a definite or fixed composition that does not vary from one sample to another.
- All substances are either *elements* or *compounds*.
- An **element** cannot be broken down into other simpler substances by chemical reactions.
  - About 100 elements known at this time
  - Each element has a *chemical symbol*: O, H, Ag, Fe, Cl, S, Hg, Au, U, etc.
- A **compound** is made up of two or more elements in *fixed* proportions, and can be broken down into simpler substances.
  - Carbon dioxide, sodium chloride, sucrose (sugar), etc.

Classifying Matter

- A **mixture** does not have a fixed composition.
- A **homogeneous mixture** has the same composition throughout, though the composition of different homogeneous mixtures may vary.
  - * Soda pop, salt water, 14K gold, and many plastics are homogeneous mixtures.
  - * 10K gold and 14K gold have different compositions but both are homogeneous.
- A **heterogeneous mixture** varies in composition and/or properties from one part of the mixture to another.
  - * Adhesive tape, CD, pen, battery, chair, and people are examples of heterogeneous mixtures.
- Most everyday "stuff" consists of mixtures.

Scientific Methods

- Scientific knowledge is *testable, reproducible, explanatory, predictive,* and *tentative*.
- In one of the most common scientific methods, we begin by constructing a **hypothesis** – a tentative explanation of the facts and observations.
- Then we design and perform **experiments** to test the hypothesis; collect **data** (measurements).
- The hypothesis is revised and the process continues.

Scientific Methods

- When our hypothesis successfully predicts what will happen, we designate it as a **scientific law** – a (usually) mathematical description of "here's what will happen."
- A **theory** is the *explanation* for a law.
  - Example: Boyle's *law* says that PV = constant for a gas sample at constant temperature.
  - Kinetic-molecular *theory* is our best explanation for Boyle's law: When atoms are squeezed into a smaller container, atoms collide more often with the walls, creating greater force and higher pressure.
- Common misconception: *theory* does not mean "imperfect fact."

Scientific Measurements

- **SI** is the International System of Units.
- In SI, there is a single base unit for each type of measurement.

Table 1.3 The Seven SI Base Units

Physical Quantity	Name of Unit	Symbol of Unit
Length	Meter*	m
Mass	Kilogram	kg
Time	Second	s
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

* Spelled *metre* in most countries other than the United States.

Scientific Measurements: SI Prefixes

- Prefixes are used to indicate powers of ten of common units that are much smaller or larger than the base unit.
- Although there are many prefixes, only a few are in very common use.
- In measurements, *kilo*-, *centi*-, and *milli*- are the three most common prefixes.

Table 1.4
Some Common SI Prefixes

Multiple	Prefix
10^{12}	*tera* (T)
10^{9}	*giga* (G)
10^{6}	*mega* (M)
10^{3}	*kilo* (k)
10^{2}	*hecto* (h)
10^{1}	*deca* (da)
10^{-1}	*deci* (d)
10^{-2}	*centi* (c)
10^{-3}	*milli* (m)
10^{-6}	*micro* (μ)*
10^{-9}	*nano* (n)
10^{-12}	*pico* (p)

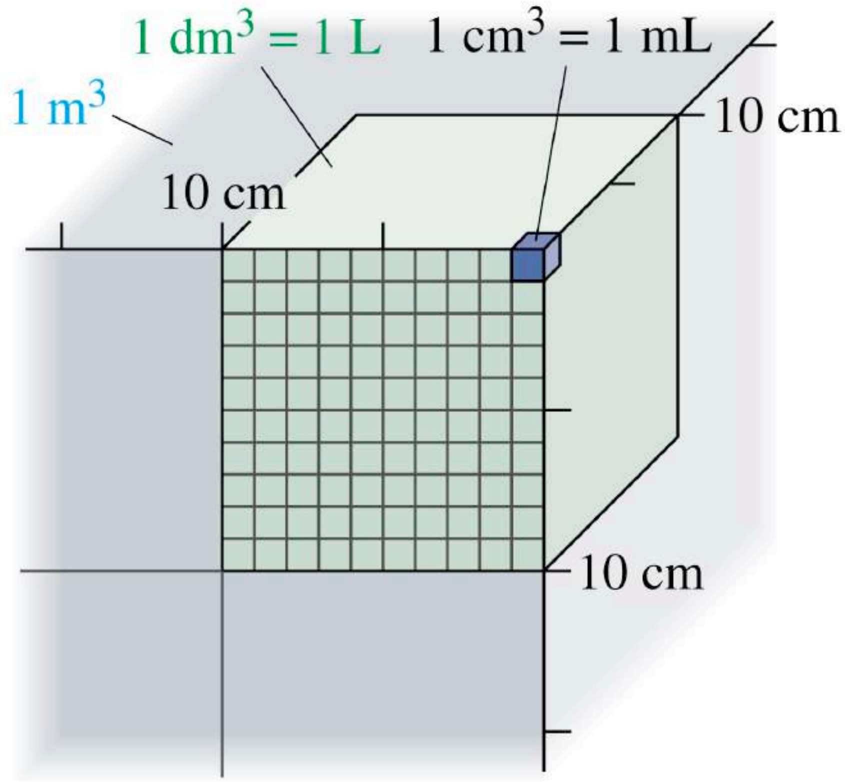
* The Greek letter μ (spelled "mu" and pronounced "mew").

Length and Area

- The base unit of length is the **meter**, a little longer than a yard.
- Common derived units include:
  - * kilometer (km; 1000 m), about 2/3 of a mile.
  - * centimeter (cm; 0.01 m) and millimeter (mm; 0.001 m)
  - * A contact lens is about 1 cm in diameter and 1 mm thick.
- The derived unit of area is the square meter (m^2) – an area one meter on a side.

Volume

- The derived unit of volume (space taken up by an object) is the cubic meter (m^3).
- A very common unit of volume, not SI but still used, is the **liter** (L).
- The milliliter (mL; 0.001 L) is also used, as is the cubic centimeter (cm^3).
- 1 mL = 1 cm^3.
- There are about five mL in one teaspoon.



Mass and Time

- **Mass** is the quantity of matter in an object; *weight* is a force.
- The base unit of mass is the **kilogram** (kg; 1000 g); it already has a prefix.
- A 1-L bottle of soft drink weighs about a kilogram.
- Commonly used mass units include the gram and the milligram (mg; 0.001 g).
- The SI base unit of time is the **second** (s).
- Smaller units of time include the millisecond (ms), microsecond (µs), and nanosecond (ns).
- Larger units of time usually are expressed in the nontraditional units of minutes, hours, days, and years.

Example 1.1

Convert the unit of each of the following measurements to a unit that replaces the power of ten by a prefix.

(a) 9.56×10^{-3} m **(b)** 1.07×10^{3} g

Example 1.2

Use exponential notation to express each of the following measurements in terms of an SI base unit.

(a) 1.42 cm **(b)** 645 μs

Example 1.1

Convert the unit of each of the following measurements to a unit that replaces the power of ten by a prefix.

(a) 9.56×10^{-3} m **(b)** 1.07×10^{3} g

Solution

Our goal is to replace each power of ten with the appropriate prefix from Table 1.4. For example, the table tells us that 10^{-3} means we should use the prefix *milli*.

(a) 10^{-3} corresponds to the prefix *milli*; 9.56 mm

(b) 10^{3} corresponds to the prefix *kilo*; 1.07 kg

Exercise 1.1A

Restate each of the following measurements by attaching an appropriate prefix to the unit to eliminate the power of ten.

(a) 2.05×10^{-6} m **(b)** 4.03×10^{3} g **(c)** 7.06×10^{-9} s **(d)** 5.15×10^{-2} m

Exercise 1.1B

Restate each of the following measurements by changing the numerical value given to a quantity with a coefficient greater than 1 and less than 10, followed by the appropriate power of ten.

(a) 6217 g **(b)** 0.0016 s **(c)** 0.0717 g **(d)** 387 m

Example 1.2

Use exponential notation to express each of the following measurements in terms of an SI base unit.

(a) 1.42 cm **(b)** 645 µs

Solution

(a) Our goal is to find the power of ten that relates the given unit to the SI base unit. Here, the letter c (for centi), used as a prefix with the base unit meter (m), means the same thing as multiplying the base unit by 10^{-2}.
$1.42 \text{ cm} = 1.42 \times 10^{-2} \text{ m}$

(b) To change microsecond to the base unit second, we need to replace the prefix *micro* by 10^{-6}. To get our answer in the conventional form of exponential notation, that is, with the coefficient of the power of ten having a value greater than 1 and less than 10, we also need to replace the coefficient 645 by 6.45×10^{2}. The result of these two changes is
$645 \text{ µs} = 645 \times 10^{-6} \text{ s} = 6.45 \times 10^{2} \times 10^{-6} \text{ s} = 6.45 \times 10^{-4} \text{ s}$

- **Exercise 1.2A**

Use exponential notation to express each of the following measurements in terms of its SI base unit.

(a) 355 µs **(b)** 1885 km **(c)** 1350 km **(d)** 425 nm

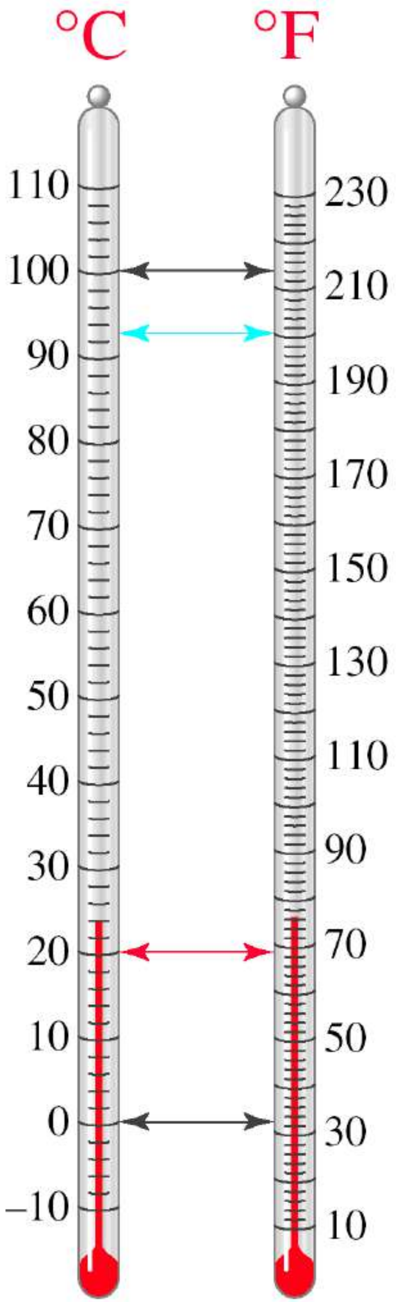
- **Exercise 1.2B**

Use exponential notation to express each of the following measurements in terms of its SI base unit.

(a) 2.28×10^{5} g **(b)** 0.083 cm **(c)** 4.05×10^{2} µm **(d)** 20.25 min

Temperature

- Temperature is the property that tells us the direction that heat will flow.
- The base unit of temperature is the **kelvin** (**K**).
- We often use the **Celsius** scale (°C) for scientific work.
  - On the Celsius scale, 0 °C is the freezing point of water, and 100 °C is the boiling point.
- The **Fahrenheit** scale (°F) is most commonly encountered in the U.S.
  - On the Fahrenheit scale, freezing and boiling water are 32 °F and 212 °F, respectively.
- $T_F = 1.8T_C + 32$
- $T_C = (T_F - 32)/1.8$



Example 1.3

At home you like to keep the thermostat at 72 °F. While traveling in Canada, you find the room thermostat calibrated in degrees Celsius. To what Celsius temperature would you need to set the thermostat to get the same temperature you enjoy at home?

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Strategy

The general approach to conversions between the Fahrenheit and Celsius temperature scales is to use the appropriate form of the equations written above. The equation for converting a Fahrenheit temperature to the Celsius scale is

$$T_C = \frac{T_F - 32}{1.8}$$

Solution

You merely need to substitute the given Fahrenheit temperature and solve for T_C:

$$T_C = \frac{72 - 32}{1.8} = \frac{40}{1.8} = 22\ °C$$

Assessment

One simple check of this calculation is to see that the numerical value of the temperature in degrees Celsius is less than the original temperature in degrees Fahrenheit, as suggested in Figure 1.7. (Note that this statement is true only for temperatures greater than –40 °F.)

Exercise 1.3A

Carry out the following temperature conversions.

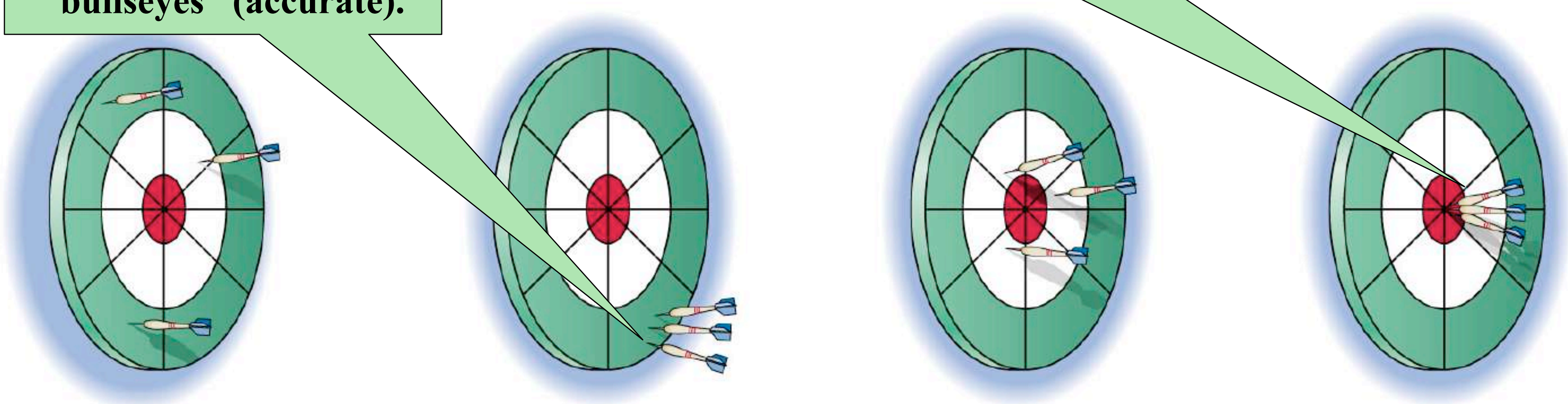
(a) 85.0 °C to °F **(b)** –12.2 °C to °F **(c)** 335 °F to °C **(d)** –20.8 °F to °C

Precision and Accuracy in Measurements

- **Precision** – how closely repeated measurements approach one another.
- **Accuracy** – closeness of measurement to "true" (accepted) value.

Darts are close together (precise) but they aren't "bullseyes" (accurate).

Darts are close together AND they are "bullseyes."



(a) Low accuracy Low precision

(b) Low accuracy High precision

(c) High accuracy Low precision

(d) High accuracy High precision

Precision and Accuracy in Measurements

- In the real world, we never know whether the measurement we make is *accurate* (why not?)
- We make repeated measurements, and strive for *precision.*
- We hope (not always correctly) that good *precision* implies good *accuracy.*

Significant Figures

- We do not want to claim more precision in our work than we actually obtained.
- ***Significant figure*** convention is observed so that the answer we report represents the *precision* of our measurements.
- Example of the concept: If you drive 273.0 miles on a fill-up of 14.1 gallons of gasoline, the calculator says that your mileage is:
  - 273.0 mi/14.1 gal = 19.36170213 mi/gal
  - Does this mean that you can predict how far your car will go on a gallon of gas – to the nearest 0.00000001 mile (about 1/1000 inch!)??
  - Of course not! – some of those digits are meaningless. (Which ones??)

Significant Figures

- **Significant figures** – all known digits, plus the first uncertain digit.
- In significant figure convention:
  - * We first determine the number of significant figures in our data (measurements).
  - * We use that knowledge to report an appropriate number of digits in our answer.
- Significant figure convention is *not* a scientific law!
- Significant figure convention is a set of *guidelines* to ensure that we don't over- or underreport the precision of results – at least not too badly…

Significant Figures in *Data*

- **Data** = measurements. (Results = calculations)
- All nonzero digits in data are significant.
- Zeroes may or may not be significant.
- To determine the number of significant figures in a measurement:
  - * Begin counting with the first nonzero digit.
  - * Stop at the end of the number.
- Problem: Zeroes in numbers without a decimal (100 mL, 5000 g) may or may not be significant.
  - * To avoid ambiguity, such numbers are often written in scientific notation:
  - * 1000 mL (?? sig fig) 1.00×10^3 mL (3 sig fig)

Significant Figures in *Data*

- *Defined* and *counting* numbers do not have uncertainty.
- ***14*** people
- ***1000*** m = ***1*** km
- *7* beakers
- The numbers 14, 1000, 1, and 7 are ***exact***.
- They have as many figures as are needed.

Significant Figures in Calculations

- General: Base the number of digits in a result on the *measurements* and not on *known values* (such as atomic masses, accurately known densities, other physical constants, etc.)
- Multiplication and division:
  - * Use the same number of *significant figures* in the result as the data with the fewest *significant figures*.
- Addition and Subtraction:
  - * Use the same number of *decimal places* in the result as the data with the fewest *decimal places*.

Example 1.4

Calculate the area, in square meters, of the poster board whose dimensions are given in Table 1.5. Report the correct number of significant figures in your answer.

Example 1.5

For a laboratory experiment, a teacher wants to divide all of a 453.6-g sample of sulfur equally among the 21 members of her class. How many grams of sulfur should each student receive?

Example 1.6

Perform the following calculation, and round off the answer to the correct number of significant figures.

$49.146\text{ m} + 72.13\text{ m} - 9.1434\text{ m} = ?$

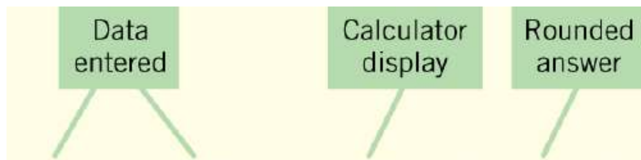
Example 1.4

Calculate the area, in square meters, of the poster board whose dimensions are given in Table 1.5. Report the correct number of significant figures in your answer.

Strategy

The area of the rectangular poster board is the product of its length and width. In expressing the result of this multiplication, we can show only as many significant figures as are found in the least precisely stated dimension: the 0.762-m width (three significant figures).

Solution



$$1.827\ \text{m} \times 0.762\ \text{m} = 1.392174\ \text{m}^2 = 1.39\ \text{m}^2$$

Assessment

When using a calculator, the display often has more digits than are significant (Figure 1.10). We use the rules for rounding off numbers as the basis for dropping the digits “2174.”

- **Exercise 1.4A**

  Calculate the volume, in cubic meters, of the poster board described in Table 1.5, given that its thickness is 6.4 mm. Report your answer to the correct number of significant figures.

- **Exercise 1.4B**

  Calculate the volume of the poster board in cubic centimeters, expressing your answer to the appropriate number of significant figures.

Example 1.5

For a laboratory experiment, a teacher wants to divide all of a 453.6-g sample of sulfur equally among the 21 members of her class. How many grams of sulfur should each student receive?

Strategy

Here we need to recognize the number “21” as a counted number. It is therefore an exact number and not subject to significant figure rules. The answer should carry four significant figures, the same as in 453.6 g.

Solution

Divide the total mass by the number of students.

$$? \text{ g} = \frac{453.6 \text{ g}}{21} = 21.60 \text{ g}$$

Assessment

In this calculation, a calculator displays the result “21.6.” We need to add the digit “0” to indicate that the result is precise to four significant figures.

- **Exercise 1.5A**

The experiment described in Example 1.5 also requires that each student have available 2.04 times as much zinc as sulfur (by mass). What total mass of zinc will the teacher need, expressed with the appropriate number of significant figures?

- **Exercise 1.5B**

Calculate the total surface area, in square centimeters, of a sugar cube that is 9.2 mm on edge, and express this area with the appropriate number of significant figures.

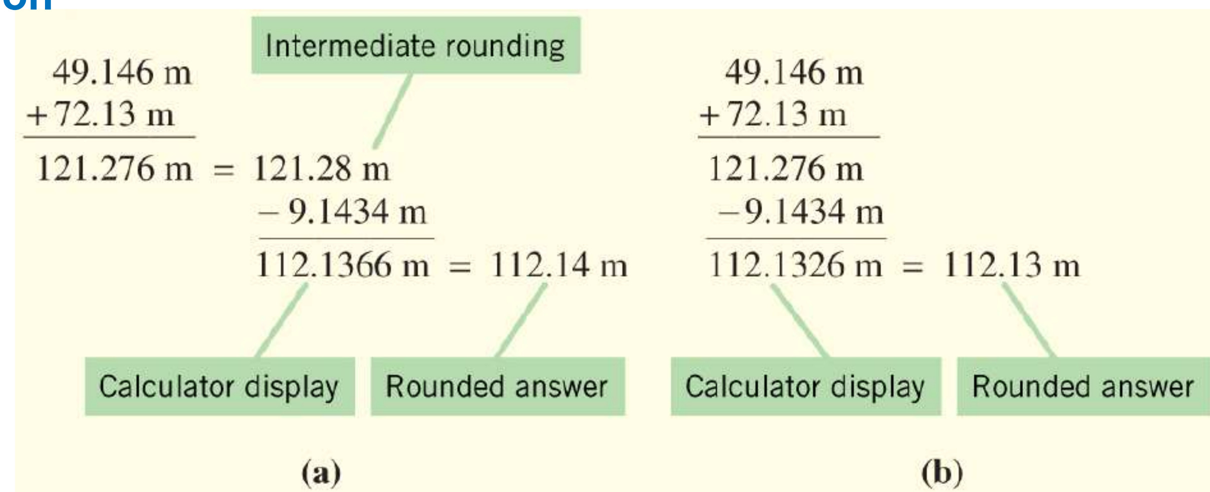
Example 1.6

Perform the following calculation, and round off the answer to the correct number of significant figures.
49.146 m + 72.13 m – 9.1434 m = ?

Strategy

We must add two numbers and then subtract a third from their sum. We do this in two ways below. In both cases, we express the answer to two decimal places, the number of decimal places in "72.13 m."

Solution



Assessment

The slight difference in the answers in (a) and (b) stems from a difference in rounding off. In method (a) we round off an intermediate result—121.276. Then we round off the final answer: 112.1366 m to 112.14 m. In method (b) we do *not* round off the intermediate result and round off only the final answer: 112.1326 m to 112.13 m. When we use a calculator, we generally do not need to write down or otherwise take note of an intermediate result.

A Problem-Solving Method

The **unit-conversion method** is based on two general concepts:

- Multiplying a quantity by *one* does not change the quantity.
- The same quantity (or *unit*) in both numerator and denominator of a fraction will *cancel*.

Unit Conversion: Conversion Factors

Table 1.6 Some Conversions Between Common (U.S.) and Metric Units

Metric		Common
Mass		
1 kg	=	2.205 lb
453.6 g	=	1 lb
28.35 g	=	1 ounce (oz)
Length		
1 m	=	39.37 in.
1 km	=	0.6214 mi
2.54 cm	=	1 in.*
Volume		
1 L	=	1.057 qt
3.785 L	=	1 gal
29.57 mL	=	1 fluid ounce (fl oz)

* U.S. inch is defined as exactly 2.54 cm. The other equivalencies are rounded off.

Example: 2.54 cm = 1 in.

We can write two conversion factors:

$$\frac{2.54\text{ cm}}{1\text{ in.}} = 1 \qquad \frac{1\text{ in.}}{2.54\text{ cm}} = 1$$

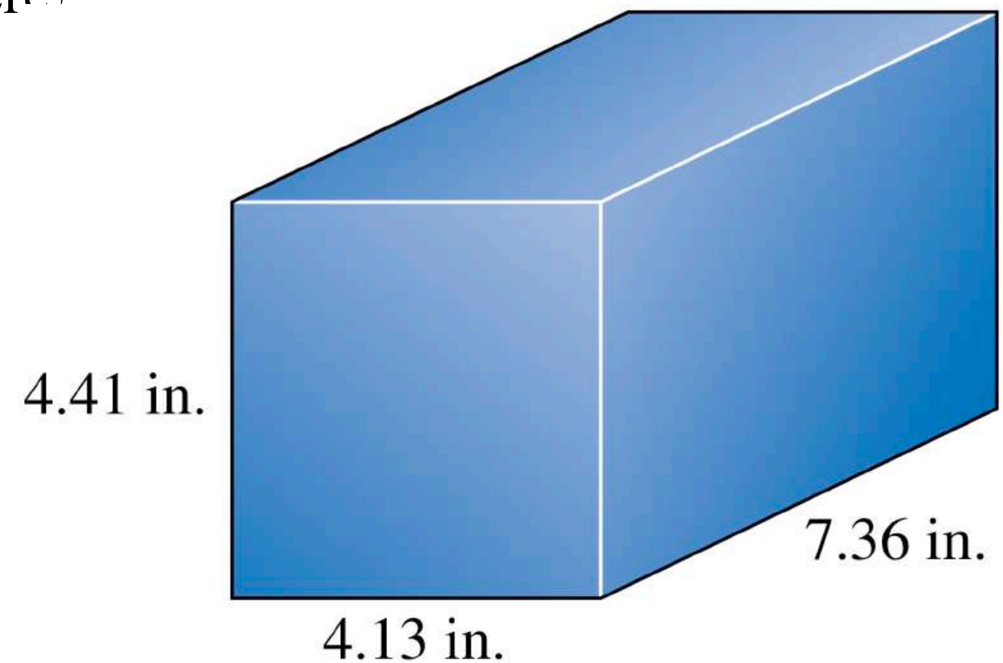
- We use these conversion factors to convert *in.* to *cm* and to convert *cm* to *in.*
- Multiply the quantity we are given by the appropriate factor.
- Question: Which factor is used for each task?
- Answer: Use the one that *cancels* the unit we do not need, and *leaves* the unit we want.

Example 1.7

What is the length in millimeters of a 1.25-ft rod?

Example 1.8

What is the volume, in cubic centimeters, of the block of wood pictured here?



Example 1.7

What is the length in millimeters of a 1.25-ft rod?

Strategy

We could make this conversion with a single conversion factor if we had a relationship between feet and millimeters. However, Table 1.6 does not provide such a relationship. We therefore need to do three successive conversions:

1. From feet to inches using the fact that 1 ft = 12 in.
2. From inches to meters with data from Table 1.6.
3. From meters to millimeters based on the prefixes in Table 1.4.

Solution

We can write three answers, one for each conversion, with the third answer being our final result. It is just as easy, however, to combine the three conversions into a single setup without writing any intermediate answers. This is the procedure sketched below.

We want the unit mm and the number that goes with it (?).

We start here.

This converts ft to in.

This converts in. to m.

We get the desired unit here.

Our answer: the number the unit

$$? \text{ mm} = 1.25 \cancel{\text{ft}} \times \frac{12 \cancel{\text{in.}}}{1 \cancel{\text{ft}}} \times \frac{1 \cancel{\text{m}}}{39.37 \cancel{\text{in.}}} \times \frac{1000 \text{ mm}}{1 \cancel{\text{m}}} = 381 \text{ mm}$$

Example 1.7 continued

Assessment

Now let's look at the use of significant figures in this problem. The measured quantity, 1.25 ft, is given with *three* significant figures. The relationship 12 in. = 1 ft is *exact*, and this will not affect the precision of our calculation. It is understood that the relationship 1 m = 39.37 in. is stated to *four* significant figures; there is no need to write 1 m as 1.000 m. Finally, the relationship 1000 mm = 1 m is *exact*; this defines the way we relate millimeters and meters. Our answer should therefore have *three* significant figures.

- **Exercise 1.7A**

Carry out the following conversions.

(a) 76.3 mm to meters **(b)** 0.0856 kg to milligrams

(c) 0.556 km to feet

- **Exercise 1.7B**

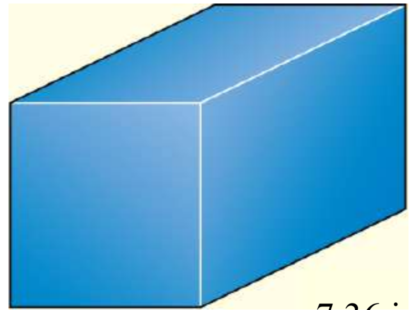
Carry out the following conversions.

(a) 1.95×10^{-3} oz to μg **(b)** 3.50 gal to fluid ounces

(c) 7.75×10^{4} ft to kilometers

Example 1.8

What is the volume, in cubic centimeters, of the block of wood pictured here?



Strategy

The volume of the block is the product of the three dimensions given in the drawing. We can proceed in one of two ways: (1) change each dimension from inches to centimeters and obtain the product in cubic centimeters or (2) obtain the volume in cubic inches and then convert to cubic centimeters.

Solution

In method (1), the factor 2.54 cm/1 in. appears three times in the setup:

$$? \text{ cm}^3 = 4.13 \cancel{\text{in.}} \times \left(\frac{2.54 \text{ cm}}{1 \cancel{\text{in.}}}\right) \times 4.41 \cancel{\text{in.}} \times \left(\frac{2.54 \text{ cm}}{1 \cancel{\text{in.}}}\right) \times 7.36 \cancel{\text{in.}} \times \left(\frac{2.54 \text{ cm}}{1 \cancel{\text{in.}}}\right)$$

$$= 2.20 \times 10^3 \text{ cm}^3$$

In method (2), the three conversion factors are combined into a single factor $(2.54 \text{ cm}/1 \text{ in.})^3$:

$$? \text{ cm}^3 = (4.13 \cancel{\text{in.}} \times 4.41 \cancel{\text{in.}} \times 7.36 \cancel{\text{in.}})^3 \times \frac{(2.54)^3 \text{ cm}^3}{(1)^3 \cancel{\text{in.}^3}} = 2.20 \times 10^3 \text{ cm}^3$$

Example 1.9

The commonly accepted measurement now used by dietary specialists in assessing whether a person is overweight is the *body mass index* (BMI), which is based on a person's mass and height. It is the mass, in kilograms, divided by the *square* of the height in meters. Thus, the units for BMI are kg/m^2. Generally speaking, if the BMI exceeds 25, a person is considered overweight. What is the BMI of a person who is 69.0 inches tall and weighs 158 lb?

Example 1.9

The commonly accepted measurement now used by dietary specialists in assessing whether a person is overweight is the *body mass index* (BMI), which is based on a person's mass and height. It is the mass, in kilograms, divided by the *square* of the height in meters. Thus, the units for BMI are kg/m^2. Generally speaking, if the BMI exceeds 25, a person is considered overweight. What is the BMI of a person who is 69.0 inches tall and weighs 158 lb?

Strategy

We must convert *two* units in the original measured quantities: from pounds to kilograms and from inches to meters (or, more specifically, from inches squared to meters squared). We will solve the problem in two ways. In the first approach, we begin by expressing the BMI in lb/in.^{2}. Then we can convert from lb to kg in the numerator and from in.^{2} to m^2 in the denominator. In an alternate approach, we do the conversions in the numerator and denominator separately and then divide the numerator by the denominator.

Solution

The conversion factors required in the first approach are

Our answer must have this unit.

We start here.

To convert lb to kg in *numerator.*

To convert in.^{2} to m^2 in *denominator.*

The answer: the number the unit

$$\text{BMI } (\text{kg/m}^2) = \frac{158\ \cancel{\text{lb}}}{(69.0)^2\ \cancel{\text{in.}^2}} \times \frac{1\ \text{kg}}{2.205\ \cancel{\text{lb}}} \times \frac{(39.37)^2\ \cancel{\text{in.}^2}}{1\ \text{m}^2} = 23.3\ \text{kg/m}^2$$

Example 1.9 continued

Solution continued

The three steps in the alternate approach are

$$\text{Numerator: } 158\,\cancel{\text{lb}} \times \frac{1\text{ kg}}{2.205\,\cancel{\text{lb}}} = 71.7\text{ kg}$$

$$\text{Denominator: } (69.0)^2\,\cancel{\text{in.}^2} \times \frac{1\text{ m}}{39.37\,\cancel{\text{in.}}} \times \frac{1\text{ m}}{39.37\,\cancel{\text{in.}}} = 3.07\text{ m}^2$$

$$\text{Division: BMI (kg/m}^2) = \frac{71.7\text{ kg}}{3.07\text{ m}^2} = 23.4\text{ kg/m}^2$$

Note:
in. × in. = in.^{2}

Assessment

Both methods are correct, even though the answers differ slightly. The source of the difference is in rounding off the intermediate results in the second method (71.655329 to 71.7 in the numerator and 3.071619 to 3.07 in the denominator). If we had stored these intermediate results in the memory of a calculator and divided the stored quantities, we would have obtained 23.328196, which rounds off to 23.3, just as in the first method.

- **Exercise 1.9A**
The pressure exerted by the atmosphere is 14.70 lb/in^2. What is this pressure, expressed in kilograms per square meter?

- **Exercise 1.9B**
What is the maximum weight, in pounds, that a person 5 ft 10.5 in. tall can maintain and have a BMI that does not exceed 25.0?

Density: A Physical Property and Conversion Factor

Density is the ratio of mass to volume:

$$d = \frac{m}{V}$$

Density can be used as a conversion factor.

For example, the density of methanol is 0.791 g/mL; therefore, there are two conversion factors, each equal to one:

$$\frac{0.791 \text{ g methanol}}{1 \text{ mL methanol}} \quad \text{and} \quad \frac{1 \text{ mL methanol}}{0.791 \text{ g methanol}}$$

Example 1.10

A beaker has a mass of 85.2 g when empty and 342.4 g when it contains 325 mL of liquid methanol. What is the density of the methanol?

Example 1.11

How many kilograms of methanol does it take to fill the 15.5-gal fuel tank of an automobile modified to run on methanol?

Example 1.10

A beaker has a mass of 85.2 g when empty and 342.4 g when it contains 325 mL of liquid methanol. What is the density of the methanol?

Strategy

We must obtain the mass of methanol as the difference in mass between the filled and empty beaker. Then we can divide the calculated mass of the methanol by its volume to obtain its density.

Solution

The mass of the methanol is

? g methanol = 342.4 g – 85.2 g = 257.2 g

The number of milliliters of methanol is given. The density is the ratio of mass to volume:

$$d = \frac{m}{V} = \frac{257.2\ \text{g}}{325\ \text{mL}} = 0.791\ \text{g/mL}$$

Assessment

We have correctly accounted for the mass of methanol; that is, it is neither of the masses stated in the problem but rather the difference between the two. We have obtained the proper unit for density, g/mL, and the numerical value is comparable to those for other liquids in Table 1.7.

- **Exercise 1.10A**
  What is the density of a 4.085-g steel ball having a 5.00-mm radius?

- **Exercise 1.10B**
  A 25.0-lb cast iron boat anchor displaces 1.62 qt of water when used. What is the density of the cast iron, in g/cm^3?

Example 1.11

How many kilograms of methanol does it take to fill the 15.5-gal fuel tank of an automobile modified to run on methanol?

Strategy

Our goal is to determine the mass of a certain volume of methanol. For this, we need to use conversion factor (a) previously given. Before using the factor, however, we must convert a volume in gallons to one in milliliters. Finally, we need to convert from grams to kilograms of methanol.

Solution

We can write all the required factors in a single setup:

$$? \text{ kg} = 15.5 \cancel{\text{gal}} \times \frac{3.785 \cancel{\text{L}}}{1 \cancel{\text{gal}}} \times \frac{1000 \text{ mL}}{1 \cancel{\text{L}}} \times \frac{0.791 \cancel{\text{g}}}{1 \text{ mL}} \times \frac{1 \text{ kg}}{1000 \cancel{\text{g}}} = 46.4 \text{ kg}$$

Assessment

When using a single setup, ensure that all unwanted units cancel.

- **Exercise 1.11A**

What is the volume in gallons of 10.0 kg of methanol (d = 0.791 g/mL)?

- **Exercise 1.11B**

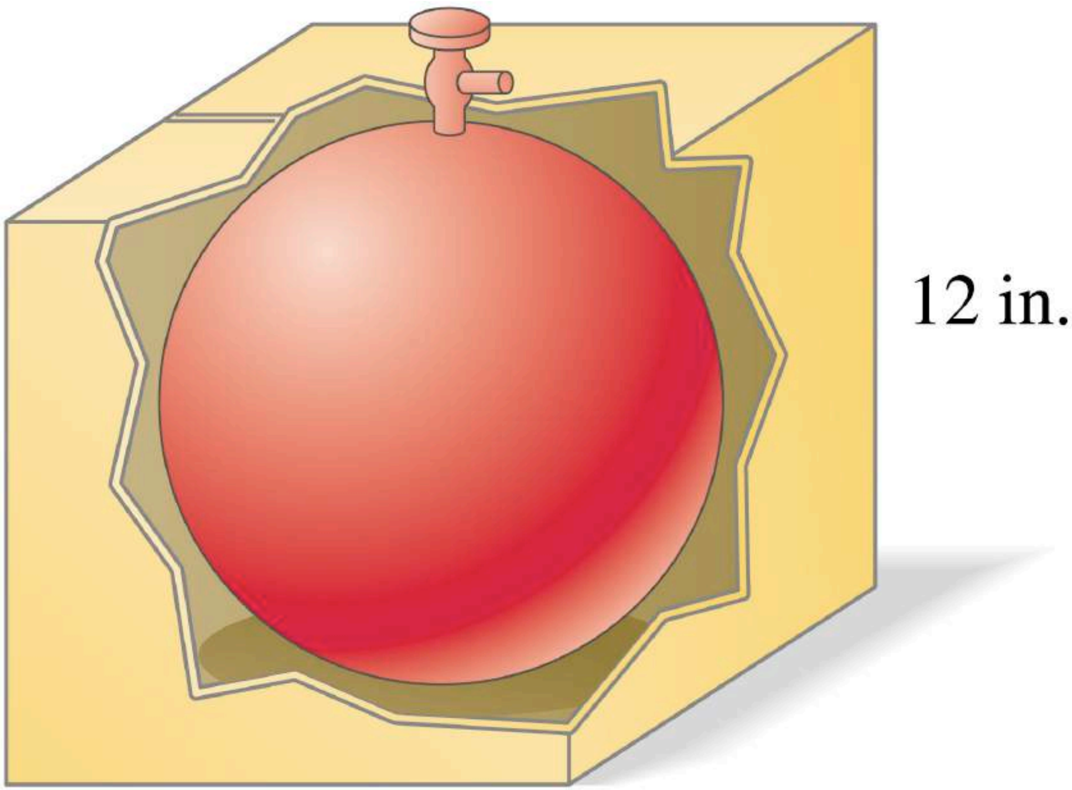
What mass of ethyl alcohol occupies the same volume as 15.0 kg of gasoline (d = 0.690 g/mL)?

Further Remarks on Problem-Solving

- A calculator may always give you an answer …
  - … but that answer is ***not always correct***.
- ***Estimation*** can be a valuable skill for determining whether an answer is correct, or for deciding among different possibilities.
- ***Estimation Examples*** and Exercises will help develop your quantitative reasoning skills.
- There is much more to science than simply plugging numbers into an equation and churning out a result on the calculator.
- To help develop your insight into chemical concepts, work the ***Conceptual Examples*** and Exercises.
- To help you integrate knowledge from several sections or chapters, work the ***Cumulative Example***.

Example 1.12 – An Estimation Example

A small storage tank for liquefied petroleum gas (LPG) appears to be spherical and to have a diameter of about 1 ft. Suppose that some common volumes for LPG tanks are 1 gal, 2 gal, 5 gal, and 10 gal. Which is the most probable volume of this particular tank?



Example 1.12—An Estimation Example

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Analysis and Conclusions

Draw a sketch similar to Figure 1.13, which shows how the tank will just fit into a cubical box about 12 in. on each side. Each edge of the box is a little over 30 cm (12 in. x 2.54 cm/1 in.). Therefore, the volume of the box is slightly larger than $V = 30\ \text{cm} \times 30\ \text{cm} \times 30\ \text{cm} = 27 \times 10^3\ \text{cm}^3 = 27\ \text{L}$. Because one gallon is slightly less than 4 L, the volume of the box is about 7 gal. The volume of the tank is clearly less than that of the box, perhaps about one-half the volume, say 3.5–4 gal. Of the volume choices given, 5 gal is the best answer.

- **Exercise 1.12A**

Which of the following is a reasonable estimate of the mass of a 20-qt pail full of water?
5 kg 10 kg 15 kg 20 kg

- **Exercise 1.12B**

With data from Example 1.12 and Figure 1.11 and a minimum of calculation, determine which of the following should have the greatest mass: (a) a 1-ft cube of balsa wood; (b) 2.00 L of water at 25 °C; (c) a 1850-g sample of liquid mercury; (d) a mixture of 1.50 L of hexane and 1.00 L of water at 25 °C.

Example 1.13 — A Conceptual Example

A sulfuric acid solution at 25 °C has a density of 1.27 g/mL. A 20.0-mL sample of this acid is measured out at 25 °C, introduced into a 50-mL flask, and allowed to cool to 21 °C. The mass of the flask plus solution is then measured at 21 °C. Use the stated data, as necessary, to calculate the mass of the acid sample.

Example 1.13—A Conceptual Example

A sulfuric acid solution at 25 °C has a density of 1.27 g/mL. A 20.0-mL sample of this acid is measured out at 25 °C, introduced into a 50-mL flask, and allowed to cool to 21 °C. The mass of the flask plus solution is then measured at 21 °C. Use the stated data, as necessary, to calculate the mass of the acid sample.

Analysis and Conclusions

The density concept tells us how to relate the mass and volume of a sample of matter. Thus, the mass of a 20.0-mL sample of the solution that is measured out at 25 °C is

$$m = d \times V = \frac{1.27\ \text{g}}{\cancel{\text{mL}}} \times 20.0\ \cancel{\text{mL}} = 25.4\ \text{g}$$

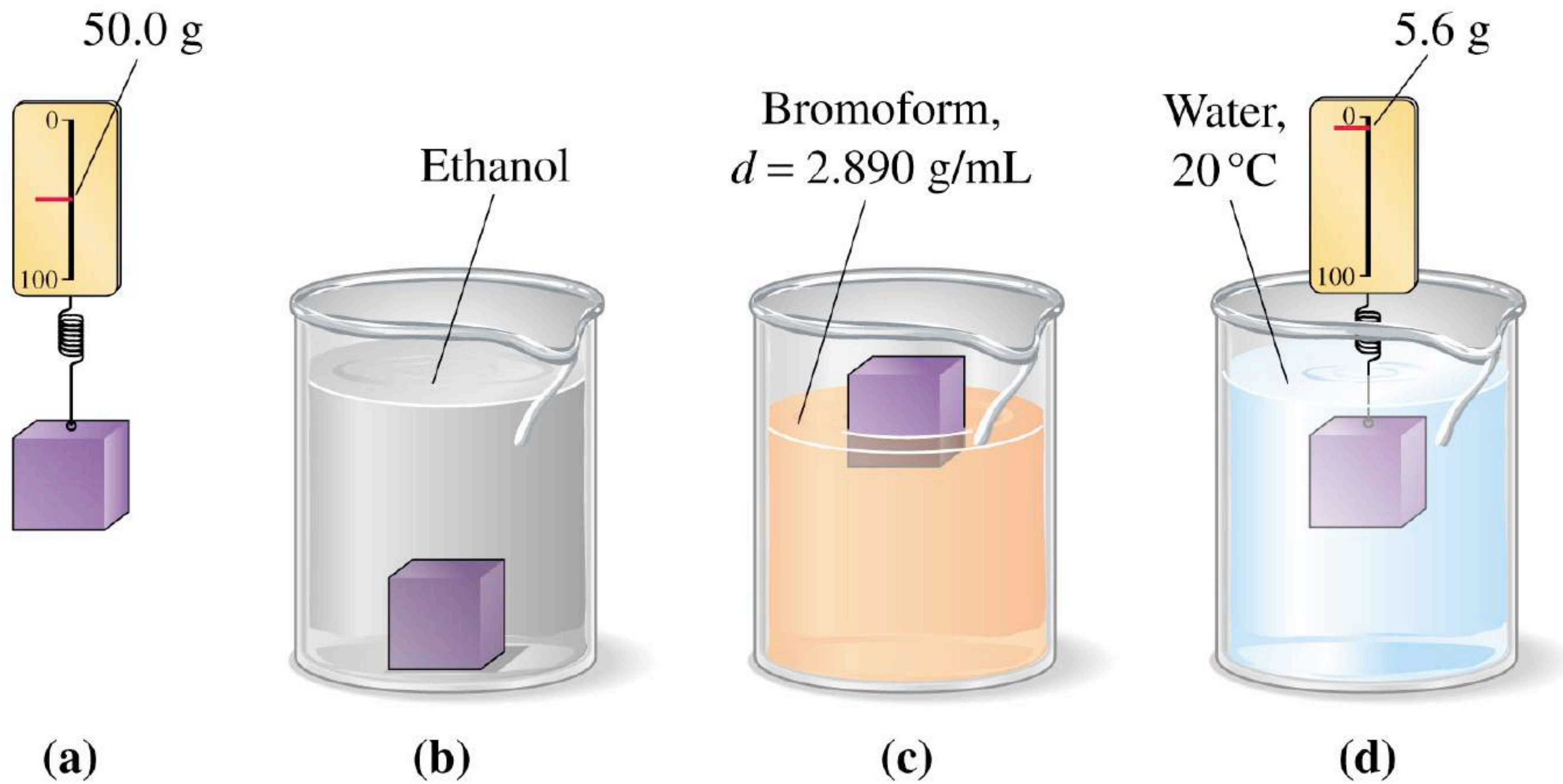
The fact that the sample is placed in a 50-mL flask is immaterial (the flask could have any volume as long as it is greater than 20 mL). Likewise, it makes no difference that the mass was measured at 21 °C because mass is not a function of temperature. That is, the sample will have a mass of 25.4 g regardless of the temperature. Finally, the fact that the solution is one of sulfuric acid is also immaterial. A 20.0-mL sample of any liquid with a density of 1.27 g/mL would also have a mass of 25.4 g under the conditions stated here.

Exercise 1.13A

In Example 1.13, why wasn't the 20.00-mL volume of solution measured at 21 °C instead of 25 °C?

Example 1.14 — A Conceptual Example

The sketches in Figure 1.14 show observations made on a small block of plastic material in four situations. What does each observation tell you about the density of the plastic?



Example 1.14—A Conceptual Example

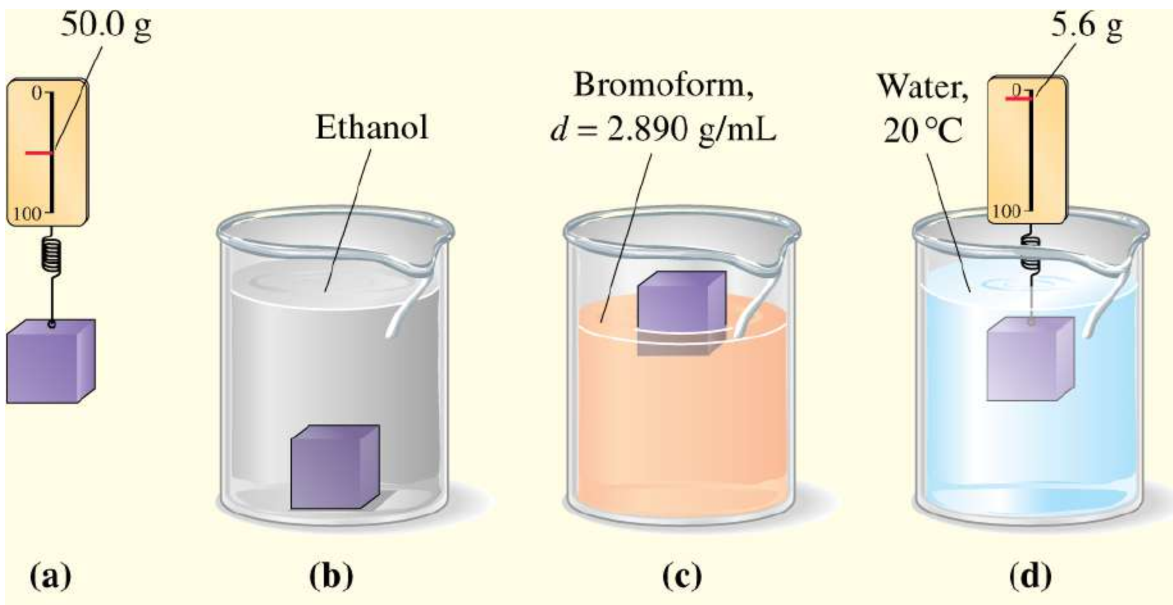
The sketches in Figure 1.14 show observations made on a small block of plastic material in four situations. What does each observation tell you about the density of the plastic?

Analysis and Conclusions

Here is what each sketch tells us:

(a) The block has a mass of 50.0 g, but this does not tell us anything about the density of the plastic material of which the block is made. For that, we also have to know the block's volume.

(b) Because the block sinks to the bottom of the ethyl alcohol, the plastic must be denser than the ethyl alcohol. That is, $d_{\text{plastic}} > 0.789$ g/mL.



▲ FIGURE 1.14 Example 1.14 visualized

Example 1.14—A Conceptual Example continued

Analysis and Conclusions continued

(c) Because the block floats on bromoform, the density of the plastic must be less than that of bromoform. That is, $d_{plastic} < 2.890$ g/mL. Moreover, because the block is about 40% submerged, the volume of bromoform having the same 50.0-g mass as the block is only about 40% of the volume of the block. Thus, using the expression $V = m/d$, we can write

$$\text{Volume of displaced bromoform} \approx 0.40 \times \text{volume of block}$$

$$\frac{\text{Mass of bromoform}}{\text{Density of bromoform}} = \frac{0.40 \times \text{mass of block}}{d_{\text{plastic}}}$$

$$\frac{50.0\text{ g}}{2.890\text{ g/mL}} = 0.40 \times \frac{50.0\text{ g plastic}}{d_{\text{plastic}}}$$

$$d_{\text{plastic}} \approx 0.40 \times 2.890\text{ g/mL} \approx 1.2\text{ g/mL}$$

(d) When submerged in water, the block displaces its volume of water. The mass of water displaced is the difference between the block's mass measured in air (from step a) and the block's mass measured in water: (50.0 – 5.6) g = 44.4 g. The volume of this mass of water = mass/density = 44.4 g/1.00 g/mL = 44.4 mL. Thus the density of the plastic is 50.0 g/44.4 mL = 1.13 g/mL.

Exercise 1.14A

Modify Figure 1.11 by removing the appropriate object and replacing it with the block of plastic described in Example 1.14, floating at the appropriate depth.

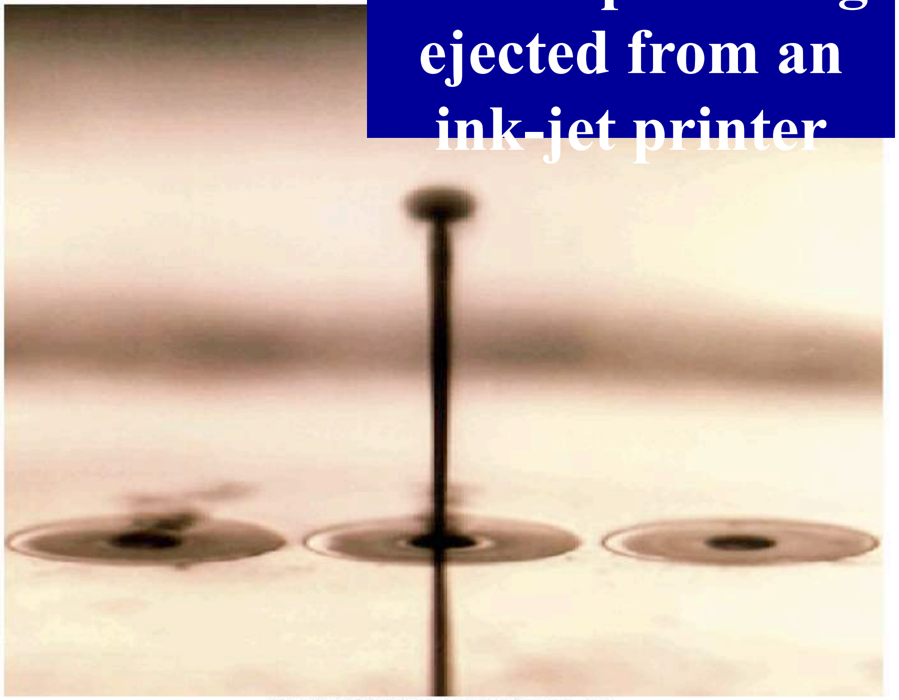
Cumulative Example

The volume of droplets generated by ink-jet printers is described in the essay "Where Smaller Is Better" on page 9.

(a) What is the diameter, in micrometers, of a spherical ink droplet from an early version of an ink-jet printer if the volume of the droplet is 200 ± 10 pL?

(b) If the ink has a density of 1.1 g/mL, what is the mass in milligrams of ink in a droplet from this printer? Is "milligrams" an appropriate unit for describing this mass?

Ink droplet being ejected from an ink-jet printer



Cumulative Example

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Strategy

In part (a), we can use the formula for the volume of a sphere, along with appropriate conversion factors, to find the dimensions (radius and then diameter) of the spherical droplet. In part (b), we can use the density of the ink as a conversion factor to find mass from the volume of the ink droplet. An additional conversion is required to obtain the desired unit of mg.

Cumulative Example continued

Solution

(a) To obtain the volume in units of $(\text{length})^3$, convert picoliters to cubic centimeters, using the equivalence of one liter to $1000\ cm^3$.

$$200\ \cancel{pL} \times \frac{10^{-12}\ \cancel{L}}{1\ \cancel{pL}} \times \frac{1000\ cm^3}{1\ \cancel{L}} = 2.0 \times 10^{-7}\ cm^3$$

To calculate the radius of the spherical drop in centimeters, use the formula for the volume of a sphere.

$$V = \frac{4\pi r^3}{3}$$

$$2.0 \times 10^{-7}\ cm^3 = \frac{4(3.14)r^3}{3}$$

$$r^3 = 3(2.0 \times 10^{-7}\ cm^3)/12.56 = 4.8 \times 10^{-8}\ cm^3$$

$$r = 3.6 \times 10^{-3}\ cm$$

Next convert the radius from centimeters to micrometers.

$$r = 3.6 \times 10^{-3}\ \cancel{cm} \times \frac{10^{-2}\ \cancel{m}}{1\ \cancel{cm}} \times \frac{1\ \mu m}{10^{-6}\ \cancel{m}} = 36\ \mu m$$

Finally, find the diameter by multiplying the radius by 2.

$$d = 2r = 72\ \mu m$$

(b) To calculate the mass of a droplet, use the volume obtained in part (a) and the density of the ink (1.1 g/mL), expressed as $1.1\ g/cm^3$.

$$2.0 \times 10^{-7}\ \cancel{cm^3} \times \frac{1.1\ g}{1\ \cancel{cm^3}} = 2.2 \times 10^{-7}\ g$$

Finally, a simple conversion gives mass in mg.

$$2.2 \times 10^{-7}\ \cancel{g} \times \frac{1\ mg}{10^{-3}\ \cancel{g}} = 2.2 \times 10^{-4}\ mg$$

Although milligrams is an acceptable unit for mass, a smaller unit might be more appropriate for this tiny mass—for example, either 0.22 μg or 220 ng rather than 2.2×10^{-4} mg. Note that whichever unit we choose, the mass is known only to two significant figures.

Assessment

Because the ink has a density of about $1\ g/cm^3$ and because $1\ cm^3 = 1000\ mm^3$, the mass of $1\ cm^3$ of ink is about 1 g, and the mass of $1\ mm^3$ of ink is about 1/1000 g = 1 mg. A droplet of ink is much smaller than $1\ mm^3$ and thus should have a mass considerably smaller than 1 mg—which it does in our answer.