General Chemistry I

Chemistry 150
Fall 2014

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Analytical Chemistry
Instrumentation
General Chemistry
Environmental Chemistry

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Please make sure that you are formally registered for Chemistry150 class (this class).

If you are not registered yet, please do the needful to register (rules and regulations) ASAP.

Chemistry150 requires your *undivided* attention to be successful. Please keep-up with class material.

Please visit with me to resolve any academic difficulties.

I keep a record of class attendance.

Chapter 1 is a general revision/introduction to some concepts in chemistry.

CHAPTER 1
Matter, Energy, and the Origins of the Universe

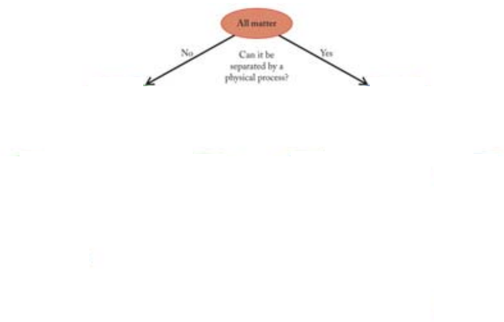
Chemistry: Definitions

- Chemistry - study of *matter*; its composition, structure, and (chemical and physical) properties.
- Matter - anything that occupies *space and has a mass.*
- Mass - defines the quantity of matter in an object.

Classes of Matter

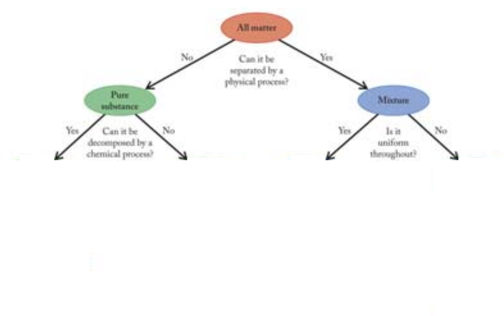
Types of Matter:

1. Pure substances
   - » Same uniform physical and chemical properties throughout the sample.
   - » Cannot be separated into simpler substances by a *physical processes (such as filtration, distillation, cooling, dissolution etc.)*.
2. Mixtures
   - » Composed of two or more substances.
   - » Can be separated by *physical processes*.



Pure Substances

- Elements:
  - A pure substance that cannot be broken down into simpler substances by *chemical or physical* means e.g. copper, zinc, carbon. (~90 common).
- Compounds:
  - Pure substance composed of two or more elements combined in definite proportions e.g. C, H and O can form; sugar $C_{12}H_{22}O_{11}$, aspirin $C_9H_8O_4$.
  - *Can be broken down into individual elements by chemical means*. (~50 million.)



An Atomic View

- Atom:
  - The smallest particle of an element that retains the chemical properties of that element.
- Molecule:
  - A collection of atoms chemically bonded together and having constant elemental proportions.
    The smallest particle of a compound that retains the chemical properties of that compound.
- Chemical (Molecular) Formula:
  - Uses symbols to identify the elements in a compound, and subscripts to indicate proportions of elements.

Molecular Formula: The notation showing the number of each type atom in the molecule.

Aspirin - molecular formula: $C_9H_8O_4$

atom ratio C:H:O = 9:8:4 (always)
mass ratio C:H:O = 108:8:64 (always)

Compounds

- Law of Constant Composition:
  - Every sample of a compound always contains the same elements in the same proportions.
  - Water (H_2O)
    - » Consists of two atoms of hydrogen (H) combined with one atom of oxygen (O).
    - » Elements and proportions represented by chemical formula (Section 1.2).

Chemical makeup of a substance points to the elements and elemental composition of the material.

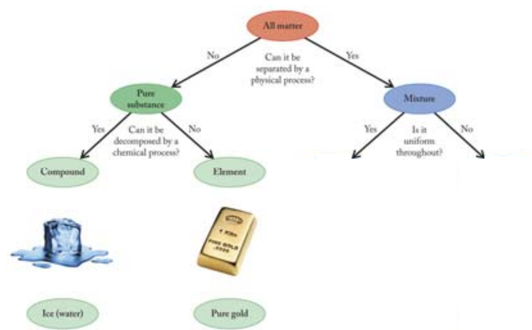
Composition of a pure substance is it's basic identity.

e.g. Water- hydrogen:oxygen = 1:8 by *weight*
Water- hydrogen:oxygen = 2:1 by *atoms*
regardless of the size of the sample of material.

(Pure) substance: Material that has a fixed elemental composition and distinct properties (they are either elements; 100% - one element, *or* compound).

Mixtures

- Homogeneous:
  - Constituents are distributed uniformly throughout the sample.
    - » Examples: salt water, brass (metal alloy, Cu+Zn)
- Heterogeneous:
  - Individual components can be seen as separate substances, non uniform properties.
    - » Examples: chocolate chip cookies, concrete



An Atomic View

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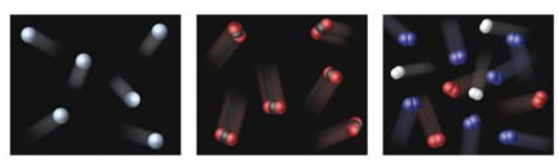
  - The smallest particle of an element that retains the chemical characteristics of that element.
- Molecule:
  - A collection of atoms chemically bonded together and having constant elemental proportions.
- Chemical Formula:
  - Uses symbols to identify the elements in a compound, and subscripts to indicate proportions of elements the compound.

Visualization: sphere = atom



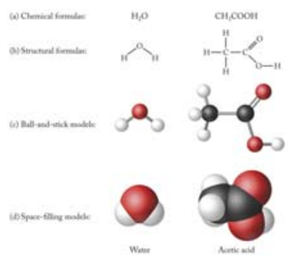
(a) Atoms of helium (b) Molecules of carbon dioxide (c) Mixture of gases

Chemistry: The Science in Context 3/e Figure 1.3
NRH Photography

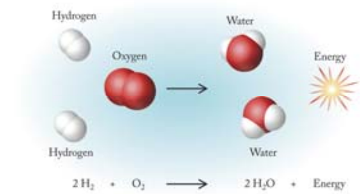
Chemical Formulas

H_2O $C_2H_4O_2$

- Chemical bonds link atoms together to make molecules.
- Chemical formulas can be represented in four ways:
  - Chemical/molecular formulas
  - Structural formulas
  - Ball-and-stick models
  - Space-filling models



Chemical Reactions



- Chemical reactions:
  - The transformation of one or more substances into different substances.
- Chemical equations:
  - Use ***chemical formulas*** to express the identities of substances involved in a reaction.
  - Use ***coefficients*** to indicate quantities of substances involved in a reaction.

Separating Mixtures

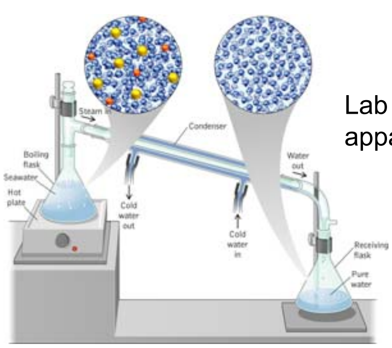
- Constituents in a mixture are isolated by physical means (*i.e.*, no chemical reactions are needed or carried out).
  - Filtration—separate a solid from a liquid by passing through a filter.
  - Distillation—separate a liquid from mixture by evaporation and re-condensation.

Filtration

a) Practical filtration

b) lab filtration

Distillation



Lab distillation apparatus

Distillation (cont.)

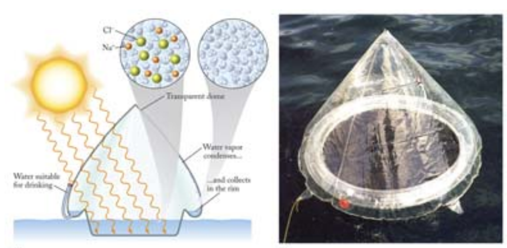


Fig. 1.12: Collection of freshwater from seawater using a solar still.

Properties of Matter

- Intensive property:
  - Independent of the amount of substance present.
  - Examples: color, hardness, density (*d* = mass/volume).
- Extensive property:
  - Varies with the quantity of the substance present.
  - Examples: volume, mass, etc.

Properties of Matter (cont.)

- Physical properties:

  Characteristics of a substance that can be observed without it changing into another substance. Examples: luster, hardness, color, etc.
- Chemical properties:

  Characteristics that can be observed only when a substance reacts with another substance. Examples: Carbonates produce a gas when added to acidic solutions.

Problem: Physical vs Chemical Properties

Which of the following properties of water are chemical and which are physical?

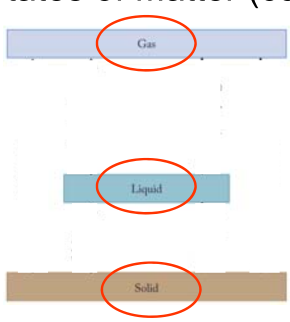
A) It normally freezes at 0.0°C.
B) It normally is useful for putting out most fires.
C) A cork floats in it, but a piece of copper sinks.
D) During digestion, starch reacts with water to form sugar.

States of Matter

- Solids:
  - Definite shapes and volumes.
- Liquids:
  - Definite volumes, indefinite shapes.
- Gases:
  - Neither definite shape nor definite volume.

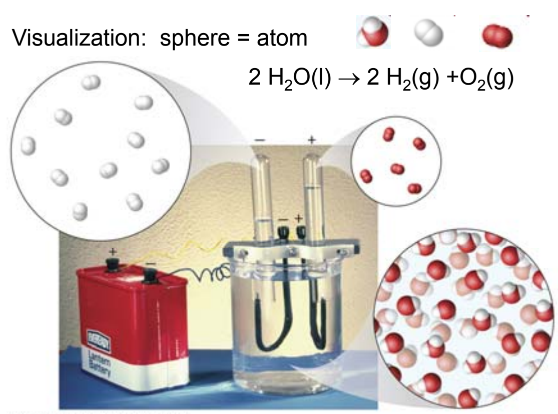
Changes of State: Transformation from one state to another due to addition or removal of heat.

States of Matter (cont.)



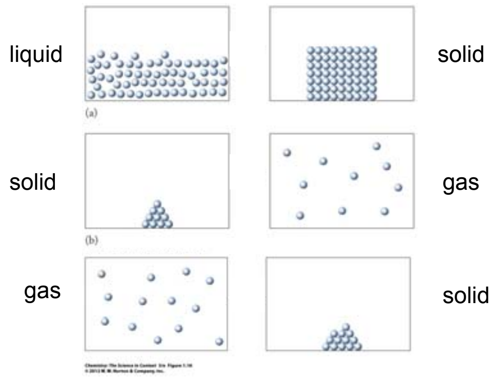
Visualization: sphere = atom

$2\ H_2O(l) \rightarrow 2\ H_2(g) + O_2(g)$



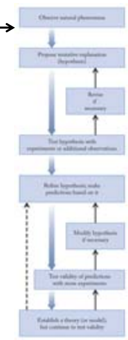
Chemistry: The Science in Context 3/e Figure 1.4
Lester V. Bergman/Corbis

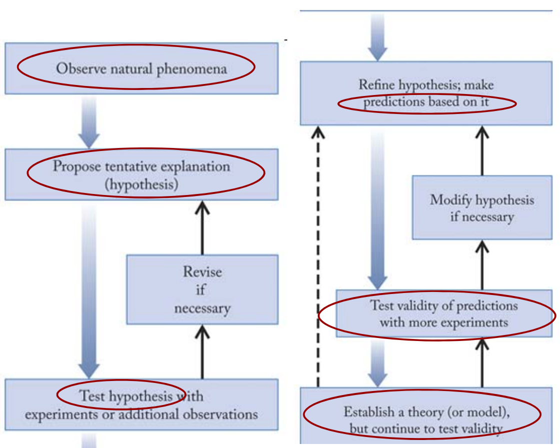
Visualization: sphere = molecule, for brevity



Scientific Method

- Scientific Method
  - Method to acquire knowledge through observation and experimentation.
- Hypothesis:
  - A tentative and testable explanation for an observation or a series of observations.
- Scientific Theory:
  - A general explanation of widely observed phenomena that have been extensively tested.





Making Measurements

- Measurements:
  - Essential for characterizing physical and chemical properties of matter. Two parts of every measurement:

5,280 feet

Number Unit

- Standardization of the units of measurement is essential.

SI Base Units

TABLE 1.2 SI Base Units

Quantity or Dimension	Unit Name	Unit Abbreviation
Mass	kilogram	kg
Length	meter	m
Temperature	kelvin	K
Time	second	s
Electric current	ampere	A
Amount of a substance	mole	mol
Luminosity	candela	cd

Prefixes for SI Units

TABLE 1.1 Commonly Used Prefixes for SI Units

PREFIX		VALUE	
Name	Symbol	Numerical	Exponential
zetta	Z	1,000,000,000,000,000,000,000	10^{21}
exa	E	1,000,000,000,000,000,000	10^{18}
peta	P	1,000,000,000,000,000	10^{15}
tera	T	1,000,000,000,000	10^{12}
giga	G	1,000,000,000	10^{9}
mega	M	1,000,000	10^{6}
kilo	k	1,000	10^{3}
hecto	h	100	10^{2}
deka	da	10	10^{1}
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}
pico	p	0.000000000001	10^{-12}
femto	f	0.000000000000001	10^{-15}
atto	a	0.000000000000000001	10^{-18}
zepto	z	0.000000000000000000001	10^{-21}

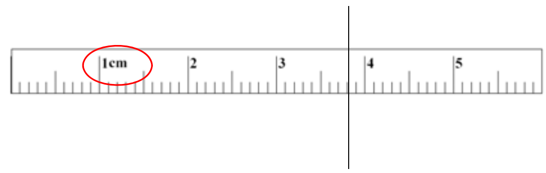
Conversions Factors are Defined Relationships and are considered exact

TABLE 1.3 Conversion Factors for SI and Other Commonly Used Units

Quantity or Dimension	Equivalent Units
Mass	1 kg = 2.205 pounds (lb); 1 lb = 0.4536 kg = 453.6 g 1 g = 0.03527 ounce (oz); 1 oz = 28.35 g
Length (distance)	1 m = 1.094 yards (yd); 1 yd = 0.9144 m (exactly) 1 m = 39.37 inches (in); 1 foot (ft) = 0.3048 m (exactly) 1 in = 2.54 cm (exactly) 1 km = 0.6214 miles (mi); 1 mi = 1.609 km
Volume	1 m^3 = 35.31 ft^3; 1 ft^3 = 0.02832 m^3 1 m^3 = 1000 liters (L) (exactly) 1 L = 0.2642 gallon (gal); 1 gal = 3.785 L 1 L = 1.057 quarts (qt); 1 qt = 0.9464 L

1 L = 1000 mL

Uncertainty in Measurements



3.82 cm

Uncertainty in Measurements

- All measurements contain an uncertainty.
  - Amount of uncertainty depends on instruments used to make the measurement.
- The digit that must be estimated is uncertain (last recorded digit).

Significant Figures

- Includes all digits known with certainty plus one digit that is uncertain.
- Rules for Counting Significant Figures:
  - Nonzero integers – count all
  - Handling zeros
    - » leading zeros – no (they are place holders)
    - » captive zeros - yes
    - » trailing zeros – yes after a decimal point
  - Exact numbers

Rules for Counting Significant Figures

- Nonzero integers—always significant.
  - 3456 → 4 sig. figs.
  - 7.35 → 3 sig. figs.
- Zeros:
  - Leading zeros not significant.
    - » 0.0392 → 3 sig. figs.

Rules for Counting Significant Figures (cont.)

- Zeros (cont.):
  - Trailing zeros not significant unless they come after a decimal point.
    - » 3700 → 2 sig. figs.
    - » 140.00 → 5 sig. figs.
  - Captive zeros always significant.
    - » 16.07 → 4 sig. figs.
- Exact numbers → infinite # of sig. figs.

Counts are exact numbers. No uncertainty.

Practice Problems

How many significant figures are in the following numbers?

- 0.04550 g = ??
- 100 lb = ??
- 101.05 mL = ??
- 350.0 g = ??

Scientific Method:
involves measurements/observations.

Observations (data)

Qualitative	Quantitative
Non numeric Data	Numeric Data (measurements)

Measurements expressed as a *number unit*.

Numbers
- Exact Numbers (no uncertainty)
  - conversion factors (defined values)
  - integral counts
- Inexact Numbers (uncertainty)
  - most measurements (real number, -with decimal)

Significant Figures in Mathematical Operations

- Multiplication/Division:
  - Number of sig. figs. in the result equals the number in with the least number of sig. figs. used in the calculation.
  - 6.380 × 2.00 = 12.76 → 12.8 (3 sig. figs.)
  - 16.84 / 2.5 = 6.736→ 6.7 (2 sig. figs.)

Significant Figures in Mathematical Operations (cont.)

- Addition / Subtraction:
  - Number of sig. figs. in the result depends on the number of decimal places of the number with the least decimal places.
  - Add, set # decimals of result that of the number with least decimals, count sig. figs.
  - 6.8 + 11.934 = 18.734 → 18.7 (3 sig. figs.)

Practice Problem

Round the answer for the mathematical operation below to the appropriate number of significant figures.

$$\frac{1.23-0.567}{0.34442}=\frac{0.663}{0.34442}=1.924978321$$

Precision and Accuracy

- Accuracy = agreement between measured value and accepted or true value.
- Precision = agreement among repeated measurements.

Which is accurate?
Which is precise?

Changing Units: Conversion Factors

- Unit conversion factor
  - a fraction in which the numerator and denominator represent equivalent quantities, but expressed in different units.
- Example
  - Equivalent quantities: 1 kg = 1000g
  - Conversion factors: $\frac{1000\text{g}}{1\text{kg}}$ and $\frac{1\text{kg}}{1000\text{g}}$

<u>Conversion between units</u>:

69.5 in = ? m

Strategy: *lose* inches
express in meters
involves multiplication(s)

starting quantity × conversion factor = desired quantity

You need to know or find: appropriate conversion factors.

<u>Relationships between different units</u> are established.
Find the conversion factors - from *<u>unit equalities</u>*.

e.g. 1 meter = 1.0936 yards (EXACTLY)
1 yard = 36 inches (EXACTLY)

69.5 in = ? Yards = ? Meters

a conversion factor $\frac{1 meter}{1.0936\, yard} = 1$ $\frac{1.0936\, yard}{1 meter} = 1$

$$69.5\text{in} \times \frac{1\text{yd}}{36\text{in}} \times \frac{1\text{m}}{1.0936\text{yd}} = 1.765321466\text{m}$$

<u>Rounding Off</u>:

Look at the left most digit to be <u>dropped</u>

<5, no change of retained digit
>5, increase retained digit by 1
=5, increase retained digit by 1, if it is odd

$$69.5\text{in} \times \frac{1\text{yd}}{36\text{in}} \times \frac{1\text{m}}{1.0936\text{yd}} = 1.765321466\text{m} = 1.76\text{m}$$

12.300 × 168.567899 = 2080.127874 (calculator)

12.300 × 168.567899 = 2080.127874 (correct)

2080.1 (correct)

12.0
3.0045
61.830452
76.834952 (calculator)

76.834952 (correct); 76.8

Exponential Notation/ scientific notation :

A number can be expressed as a multiple of 10.

$120.0 = 1.200 \times 10^{2}$
$12456.23 = 1.245623 \times 10^{4}$
$0.0187 = 1.87 \times 10^{-2}$

The exponent of (10^{x}) is considered to be exact.

Significant figures in scientific notation:

Same value but different number of significant figures.

E.g. 204100.0 g (7) = 2.041000×10^{5}g

2.041×10^{5} g (4)
2.0410×10^{5} g (5)
2.04100×10^{5} g (6)

Report to the correct number of significant figures?

$$\frac{(2.5\times10^{-3})(376-14.1)}{7.663}$$

Conversions Factors are Defined Relationships and are considered exact

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1 L = 1000 mL

Practice: Unit Conversion

1. Change 18.0 mL to liters.
2. Express 2.63 pounds in milligrams.
3. Express a volume of 1.250 L in m^3.

Density

Density = mass of substance per unit volume of the substance:

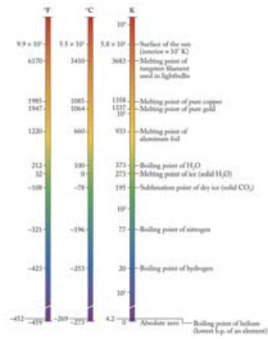
(Density: intensive or extensive property?)

Practice: Using Density

The density of Ti is 4.50 g/cm^3. What is the volume of 7.20 g of Titanium?

Temperature Scales

- Fahrenheit (°F)
  Celsius (°C)
- Kelvin (K)
- Temperature Conversions:
  - K = °C + 273.15
  - °C = 5/9 (°F − 32)



Practice: Temperature Conversions

The lowest temperature measured on the Earth is −128.6°F, recorded at Vostok, Antarctica, in July 1983. What is this temperature in °C and in Kelvin?

Sample Exercise 1.1

Which of the following properties of gold are chemical and which are physical?

a. Gold metal, which is insoluble in water, can be made soluble by reacting it with a mixture of nitric and hydrochloric acids known as aqua regia.

b. Gold melts at 1064°C.

c. Gold can be hammered into sheets so thin that light passes through them.

d. Gold metal can be recovered from gold ore by treating the ore with a solution containing cyanide, which reacts with and dissolves gold.

Sample Exercise 1.2

Which physical state is represented in each box? (The particles could be atoms or molecules.) What changes of state are indicated by the two arrows? What would the changes of state be if both arrows pointed in the opposite direction?

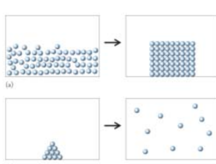


FIGURE 1.15 Changes in the physical state of matter.

Sample Exercise 1.2 (cont.)

The particles in the box on the left in Figure 1.15(b) represent a solid because they are ordered and do not adopt the shape of the box. Those in the right box are dispersed throughout the box and represent a gas. The arrow represents a solid turning into a vapor, and the state change is sublimation. The reverse process—a vapor becoming a solid—is deposition.

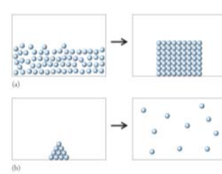


FIGURE 1.15 Changes in the physical state of matter.

Sample Exercise 1.3

A nugget of a shiny yellow mineral has a mass of 30.01 g. Its volume is determined by placing it in a 100 mL graduated cylinder containing 56.3 mL of water. The volume after the nugget is added is 62.6 mL. Is the nugget made of gold?

Sample Exercise 1.3 (cont.)

$$d = \frac{m}{V}$$
$$= \frac{30.01\ \text{g}}{(62.6 - 56.3)\ \text{mL}} = \frac{30.01\ \text{g}}{6.3\ \text{mL}} = \frac{4.7635\ \text{g}}{\text{mL}}$$

Because we know the volume of the nugget (6.3 mL) to two significant figures, we can know its density to only two significant figures, so we must round off the result to 4.8 g/mL. The density of gold is 19.3 g/mL, so the nugget cannot be pure gold.

Sample Exercise 1.4

Which of the following numerical values associated with the Washington Monument in Washington, DC, are exact numbers and which are not exact?

(a) the monument is made of 36,491 white marble blocks;
(b) the monument is 169 m tall;
(c) there are 893 steps to the top;
(d) the mass of the aluminum capstone is 2.8 kg;
(e) the area of the foundation is 1487 m^2

Sample Exercise 1.5

The Star of Africa (Figure 1.23) is one of the world's largest diamonds, having a mass of 106.04 g. What is its mass in milligrams and in kilograms?

Sample Exercise 1.5 (cont.)

- **Analyze:** We need two conversion factors, one for converting grams to milligrams and one for converting grams to kilograms. Table 1.1 tells us that 1 mg × 0.001 g and 1 kg × 1000 g. Therefore our possibilities for conversion factors are

$$\frac{1\ \text{mg}}{0.001\ \text{g}} \quad \frac{0.001\ \text{g}}{1\ \text{mg}} \quad \frac{1\ \text{kg}}{1000\ \text{g}} \quad \frac{1000\ \text{g}}{1\ \text{kg}}$$

Because we want factors showing our desired units in the numerator, we use the first and third possibilities.

Sample Exercise 1.5 (cont.)

- **Solve:** We multiply the given mass by each conversion factor to obtain the desired results:

$$106.04\ \cancel{\text{g}} \times \frac{1\ \text{mg}}{0.001\ \cancel{\text{g}}} = 106{,}040\ \text{mg} \quad 106.04\ \cancel{\text{g}} \times \frac{1\ \text{kg}}{1000\ \cancel{\text{g}}} = 0.10604\ \text{kg}$$

- **Think about It:** The results make sense because there should be many more of the smaller units (106,040 mg versus 106.04 g) in a given mass and many fewer of the larger units (0.10604 kg versus 106.04 g).

Sample Exercise 1.8

The temperature of interstellar space is 2.73 K. What is this temperature on the Celsius scale and on the Fahrenheit scale?

Sample Exercise 1.8 (cont.)

- To convert from kelvins to degrees Celsius, we have

$$\begin{aligned} K &= {}^\circ C + 273.15 \\ {}^\circ C &= K - 273.15 \\ &= 2.73 - 273.15 = -270.42^\circ C \end{aligned}$$

To convert degrees Celsius to degrees Fahrenheit, we rearrange Equation 1.2

$$^\circ C = \frac{5}{9}(^\circ F - 32)$$

to solve for degrees Fahrenheit. Multiplying both sides by 9 and dividing both by 5 gives us

$$\frac{9}{5}{}^\circ C = {}^\circ F - 32$$

$$^\circ F = \frac{9}{5}{}^\circ C + 32 = \frac{9}{5}(-270.42) + 32 = -454.76^\circ F$$

The value 32°F is considered a definition and so is not used in determining the number of significant figures in the answer. The number that determines the accuracy to which we can know this value is -270.42°C.