Chapter 1
Chemistry: The Science of Change

Practice Problems C

1.1 (iii)

1.2 (i) and (i)

1.3 pink liquid = grey solid < blue solid < yellow liquid < blue liquid < green solid

1.4 12 blue cubes, infinite number of significant figures;
2×10^1 red spheres, one significant figure

1.5 2.5×10^2 lb

1.6 2.67 g/cm^3

1.7 (a) 4 red blocks/1 object
(b) 1 object/1 yellow block
(c) 2 white blocks/1 yellow block
(d) 1 yellow block/6 grey connectors

1.8 375 red bars; 3.5×10^3 yellow balls

Key Skills

1.1 d   1.2 e   1.3 a   1.4 b

Questions and Problems

1.1 **Chemistry is the study of matter and the changes that matter undergoes.**

Matter is anything that has mass and occupies space.

1.2 **The scientific method is a set of guidelines used by scientists to add their experimental results to the larger body of knowledge in a given field. The process involves observation, hypothesis, experimentation, theory development, and further experimentation.**

1.3 **A hypothesis explains observations. A theory explains observations and laws from accumulated experiments and predicts related phenomena.**

1.4 a. **Theory** – Cell theory (biology).

b. **Hypothesis** – It's possible that this statement is true, but we would need data to substantiate it.

c. **Law** – Newton's first law of motion (physics)

1.5 a. **Law** – First law of thermodynamics

b. **Hypothesis** – This actually has not been supported by experimental evidence.

c. **Theory** – Theory of plate tectonics (geology).

1.6 a. Chemistry units: **meter (m), centimeter (cm), millimeter (mm)**

SI base unit: **meter (m)**

b. Chemistry units: **cubic decimeter (dm^3) or liter (L), milliliter (mL), cubic centimeter (cm^3)**

SI base unit: **cubic meter (m^3)**

c. Chemistry units: **gram (g)**

SI base unit: **kilogram (kg)**

d. Chemistry units: **second (s)**

SI base unit: **second (s)**

e. Chemistry units: **kelvin (K) or degrees Celsius (°C)**

SI base unit: **kelvin (K)**

1.7 a. $\mathbf{1 \times 10^{6}}$ c. $\mathbf{1 \times 10^{-1}}$ e. $\mathbf{1 \times 10^{-3}}$ g. $\mathbf{1 \times 10^{-9}}$

b. $\mathbf{1 \times 10^{3}}$ d. $\mathbf{1 \times 10^{-2}}$ f. $\mathbf{1 \times 10^{-6}}$ h. $\mathbf{1 \times 10^{-12}}$

1.8 For liquids and solids, chemists normally use **g/mL or g/cm^3** as units for density.

For gases, chemists normally use **g/L** as units for density. Gas densities are generally very low, so the smaller unit of g/L is typically used. 1 g/L = 0.001 g/mL.

1.9 Mass is a measure of the amount of matter in an object or sample. It remains constant regardless of where it is measured.

Weight is the force exerted by an object or sample due to gravity. It depends on the gravitational force where the weight is measured.

Since gravity on the moon is about one-sixth that on Earth,

$$\text{Weight on the moon } = (3495 \text{ lb on Earth})\left(\frac{1}{6}\right) = \mathbf{582.5\ lb}$$

1.10 **Kelvin is known as the absolute temperature scale, meaning the theoretically lowest possible temperature is 0 K.**

The units of the Celsius and Kelvin scales are equal in magnitude, so conversion between units is a matter of addition:

$$\mathbf{K = {}^\circ C + 273.15}$$

The freezing point of water is defined as 0°C. The boiling point of water is defined as 100°C.

In the Fahrenheit scale, the freezing point of water is 32°F, and the boiling point of water is 212°F. Since the difference is 180°F, compared to 100°C between the freezing and boiling points of water, one degree Fahrenheit represents a smaller change in temperature than one degree Celsius. To convert between these two temperature scales, use:

$$\textbf{temperature in degrees Celsius} = (\textbf{temperature in °F} - \textbf{32°F}) \times \frac{\textbf{5°C}}{\textbf{9°F}}$$

1.11 **Strategy:** Use the density equation:

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

Solution:

$$d = \frac{m}{V} = \frac{546 \text{ g}}{175 \text{ mL}} = \mathbf{3.12\ g/mL}$$

1.12 Mass of methanol $= \dfrac{0.792\ \text{g}}{1\ \text{mL}} \times 25.0\ \text{mL} = \mathbf{19.8\ g}$

1.13 **Strategy:** Find the appropriate equations for converting Fahrenheit to Celsius and Celsius to Fahrenheit given in Section 1.4 of the text. Substitute the temperature values given in the problem into the appropriate equation.

Setup: Conversion from °F to °C:

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

Conversion from °C to °F:

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

Solution: a. $°\text{C} = (95°\text{F} - 32°\text{F}) \times \dfrac{5°\text{C}}{9°\text{F}} = \mathbf{35°C}$

b. $°\text{C} = (12°\text{F} - 32°\text{F}) \times \dfrac{5°\text{C}}{9°\text{F}} = \mathbf{-11°C}$

c. $°\text{C} = (103°\text{F} - 32°\text{F}) \times \dfrac{5°\text{C}}{9°\text{F}} = \mathbf{39°C}$

d. $°\text{C} = (1852°\text{F} - 32°\text{F}) \times \dfrac{5°\text{C}}{9°\text{F}} = \mathbf{1011°C}$

e. $°\text{F} = \left(-273.15°\text{C} \times \dfrac{9°\text{F}}{5°\text{C}}\right) + 32°\text{F} = \mathbf{-459.67°F}$

1.14 a. $°\text{C} = (105°\text{F} - 32°\text{F}) \times \dfrac{5°\text{C}}{9°\text{F}} = \mathbf{41°C}$

b. $°\text{F} = \left(-11.5°\text{C} \times \dfrac{9°\text{F}}{5°\text{C}}\right) + 32°\text{F} = \mathbf{11.3°F}$

c. $°\text{F} = \left(6.3 \times 10^{3}°\text{C} \times \dfrac{9°\text{F}}{5°\text{C}}\right) + 32°\text{F} = \mathbf{1.1 \times 10^{4}°F}$

1.15 **Strategy:** Use the density equation.

Solution:

$$\text{Volume of water} = V = \frac{m}{d} = \frac{4.50\ \text{g}}{0.998\ \text{g/mL}} = \mathbf{4.51\ mL}$$

1.16 Volume of platinum $= \dfrac{90.3\text{ g}}{21.5\text{ g/cm}^3} = \mathbf{4.20\ cm^3}$

1.17 **Strategy:** Use the equation for converting °C to K.

Setup: Conversion from °C to K:

$$K = °C + 273.15$$

Solution: a. K = 97.79°C + 273.15 = **370.94 K**

b. K = 38°C + 273 = **311 K**

c. K = –39°C + 273 = **234 K**

Note that when there are no digits to the right of the decimal point in the original temperature, we use 273 instead of 273.15.

1.18 a. °C = K – 273.15 = 90.19 K – 273.15 = **–182.96°C**

b. °C = 87.30 K – 273.15 = **–185.85°C**

c. °C = 1211 K – 273 = **938°C**

1.19 Mass is extensive and additive: 18.5 + 45.8 = **64.3 g**
Temperature is intensive: **20°C**
Density is intensive: **11.35 g/cm^3**

1.20 Mass is extensive and additive: 61.1 + 95.3 = **156.4 g**
Temperature is intensive: **10°C**
Density is intensive: **1.00 g/mL**

1.21 a. **Exact.** The number of tickets is an exact number determined by counting.

b. **Inexact.** The volume is an inexact number that must be measured.

c. **Exact.** The number of eggs is an exact number determined by counting.

d. **Inexact.** The mass of oxygen is an inexact number that must be measured.

e. **Inexact.** The number of days is an inexact number that must be measured.

1.22 **Significant figures are the meaningful digits in a reported number. They indicate the level of uncertainty in a measurement. Using too many significant figures implies a greater certainty in a measured or calculated number than is realistic.**

1.23 **Accuracy tells us how close a measurement is to the true value. Precision tells us how close a series of replicate measurements are to one another. Having precise measurements does not always guarantee an accurate result, because there may be an error made that is common to all the measurements.**

1.24 a. The decimal point must be moved eight places to the right, making the exponent –8.

$$0.000000027 = \mathbf{2.7 \times 10^{-8}}$$

b. The decimal point must be moved two places to the left, making the exponent 2.

$$356 = \mathbf{3.56 \times 10^{2}}$$

c. The decimal point must be moved four places to the left, making the exponent 4.

$$47{,}764 = \mathbf{4.7764 \times 10^{4}}$$

d. The decimal point must be moved two places to the right, making the exponent –2.

$$0.096 = \mathbf{9.6 \times 10^{-2}}$$

1.25 **Strategy:** To convert an exponential number $N \times 10^n$ to a decimal number, move the decimal n places to the left if $n < 0$, or move it n places to the right if $n > 0$. While shifting the decimal, add place-holding zeros as needed.

Solution: a. 1.52×10^{-2} = **0.0152**

b. 7.78×10^{-8} = **0.0000000778**

c. 3.29×10^{-6} = **0.00000329**

d. 8.41×10^{-1} = **0.841**

1.26 a. $145.75 + \left(2.3 \times 10^{-1}\right) = 145.75 + 0.23 = 145.98 = \mathbf{1.4598 \times 10^2}$

b. $\dfrac{79{,}500}{2.5\times10^2} = \dfrac{7.95\times10^4}{2.5\times10^2} = \mathbf{3.2 \times 10^2}$

c. $\left(7.0\times10^{-3}\right) - \left(8.0\times10^{-4}\right) = \left(7.0\times10^{-3}\right) - \left(0.80\times10^{-3}\right) = \mathbf{6.2 \times 10^{-3}}$

d. $\left(1.0\times10^{4}\right) \times \left(9.9\times10^{6}\right) = \mathbf{9.9 \times 10^{10}}$

1.27 a. Addition using scientific notation

Strategy: A measurement is in *scientific notation* when it is written in the form $N \times 10^n$, where $0 \le N < 10$ and n is an integer. When adding measurements that are written in scientific notation, rewrite the quantities so that they share a common exponent. To get the "N part" of the result, we simply add the "N parts" of the rewritten numbers. To get the exponent of the result, we simply set it equal to the common exponent. Finally, if need be, we rewrite the result so that its value of N satisfies $0 \le N < 10$.

Solution: Rewrite the quantities so that they have a common exponent. In this case, choose the common exponent $n = -3$.

$$0.0095 = 9.5 \times 10^{-3}$$

Add the "N parts" of the rewritten numbers and set the exponent of the result equal to the common exponent.

$$\begin{array}{r} 9.5\times10^{-3} \\ +\ 8.5\times10^{-3} \\ \hline 18.0\times10^{-3} \end{array}$$

Rewrite the number so that it is in scientific notation (so that $0 \le N < 10$).

$$18.0 \times 10^{-3} = \mathbf{1.8 \times 10^{-2}}$$

b. Division using scientific notation

Strategy: When dividing two numbers using scientific notation, divide the "N parts" of the numbers in the usual way. To find the exponent of the result, *subtract* the exponent of the divisor from that of the dividend.

Solution: Make sure that all numbers are expressed in scientific notation.

$$653 = 6.53 \times 10^2$$

Divide the "*N* parts" of the numbers in the usual way.

$$6.53 \div 5.75 = 1.14$$

Subtract the exponents.

$$1.14 \times 10^{+2-(-8)} = 1.14 \times 10^{+2+8} = \mathbf{1.14 \times 10^{10}}$$

c. Subtraction using scientific notation

Strategy: When subtracting two measurements that are written in scientific notation, rewrite the quantities so that they share a common exponent. To get the "*N* part" of the result, we simply subtract the "*N* parts" of the rewritten numbers. To get the exponent of the result, we simply set it equal to the common exponent. Finally, if need be, we rewrite the result so that its value of *N* satisfies $0 \le N < 10$.

Solution: Rewrite the quantities so that they have a common exponent. Rewrite 850,000 in such a way that $n = 5$.

$$850{,}000 = 8.5 \times 10^5$$

Subtract the "*N* parts" of the numbers and set the exponent of the result equal to the common exponent.

$$\begin{array}{r} 8.5\times10^5 \\ -\ 9.0\times10^5 \\ \hline -\ 0.5\times10^5 \end{array}$$

Rewrite the number so that $0 \le N < 10$ (ignore the sign of *N* when it is negative).

$$-0.5 \times 10^5 = \mathbf{-5 \times 10^4}$$

d. Multiplication using scientific notation

Strategy: When multiplying two numbers using scientific notation, multiply the "*N* parts" of the numbers in the usual way. To find the exponent of the result, *add* the exponents of the two measurements.

Solution: Multiply the "*N* parts" of the numbers in the usual way.

$$3.6 \times 3.6 = 12.96$$

Add the exponents.

$$12.96 \times 10^{-4+(+6)} = 12.96 \times 10^{2}$$

Rewrite the number so that it is in scientific notation (so that $0 \leq N < 10$). Round the final result to two significant figures.

$$\mathbf{13 \times 10^{2} = 1.3 \times 10^{3}}$$

1.28 a. **four** d. **two, three, or four** g. **one**

b. **two** e. **three** h. **two**

c. **five** f. **one**

1.29 a. **one** d. **four** g. **one or two**

b. **three** e. **three** h. **three**

c. **three** f. **one**

1.30 a. **16.5 m** b. **3 kg** c. **5.11 cm^3**

1.31 a. Division

Strategy: The number of significant figures in the answer is determined by the original number having the smallest number of significant figures.

Solution:

$$\frac{19.50 \text{ km}}{6.48 \text{ km}} = 3.009$$

Because the original number 6.48 km has only three significant figures, the result of the calculation can have only three significant figures. Thus, to the appropriate number of significant figures, the answer is:

3.01

b. Subtraction

Strategy: The number of significant figures to the right of the decimal point in the answer is determined by the lowest number of digits to the right of the decimal point in any of the original numbers.

Solution: Writing both numbers in the decimal notation, we have

$$\begin{array}{r} 0.00125\ \text{mg} \\ -0.0000901\ \text{mg} \\ \hline 0.0011\mathbf{599}\ \text{mg} \end{array}$$

The bold numbers are nonsignificant digits because the number 0.00125 has only five digits to the right of the decimal point. Therefore, we carry five digits to the right of the decimal point in our answer.

The correct answer rounded off to the appropriate number of significant figures is:

$$\mathbf{0.00116\ mg = 1.16 \times 10^{-3}\ mg}$$

c. Addition

Strategy: The number of significant figures to the right of the decimal point in the answer is determined by the lowest number of digits to the right of the decimal point in any of the original numbers.

Solution: Writing both numbers with exponents = +7, we have

$$(0.911 \times 10^7\ \text{dm}) + (2.05 \times 10^7\ \text{dm}) = \mathbf{2.96 \times 10^7\ dm}$$

Since 2.05×10^7 has only two digits to the right of the decimal point, two digits are carried to the right of the decimal point in the final answer.

1.32 **Student A's results are neither precise nor accurate. Student B's results are both precise and accurate. Student C's results are precise but not accurate.**

1.33 **Carpenter Z's measurements are the most accurate. Carpenter Y's measurements are the least accurate. Carpenter X's measurements are the most precise. Carpenter Y's measurements are the least precise.**

1.34 a. The volume of liquid is **32.5 mL**. Since the graduated cylinder is marked at every 1 mL, the uncertainty of the measurement is in the tenths place. Remember to read the volume at the *bottom* of the meniscus.

b. The length of the box is **5.5 in**. Since the ruler is marked at every 1 in, the uncertainty of the measurement is in the tenths place.

1.35 **Upper thermometer: 41.85°**

Lower thermometer: 41.85°

Since both thermometers are marked to tenths of a degree, the number of significant figures is the same, regardless of whether or not the tenths of a degree are labeled.

1.36 **Strategy:** Calculate the volume of the rectangular solid using: $V = l \times w \times h$. Then, use the density equation, $d = \frac{m}{V}$, to find the mass.

Setup: Solve the density equation for m to get:

$$m = dV$$

Solution:

$$V = (2.18\text{ cm})\,(4.09\text{ cm})\,(14.25\text{ cm}) = 127\text{ cm}^3$$

$$m = dV = \frac{8.16\text{ g}}{1\text{ cm}^3} \times 127\text{ cm}^3 = \mathbf{1.04 \times 10^3\ g}$$

1.37 **Strategy:** The difference between the masses of the empty and filled bulbs is the mass of the gas in the bulb. The density is then determined using the density equation, $d = \frac{m}{V}$.

Solution: 243.22 g – 243.07 g = 0.15 g gas

$$d = \frac{m}{V} = \frac{0.15\text{ g}}{135.6\text{ mL}} = 1.1 \times 10^{-3}\text{ g/mL}$$

$$\frac{1.1 \times 10^{-3}\text{ g}}{1\text{ mL}} \times \frac{1 \times 10^{3}\text{ mL}}{\text{L}} = \mathbf{1.1\,g/L}$$

Because the density of gases is generally low, the density is typically expressed in g/L.

1.38 a. $22.6\text{ m} \times \frac{1\text{ dm}}{0.1\text{ m}} = \mathbf{226\,dm}$

b. $25.4\text{ mg} \times \frac{1 \times 10^{-3}\text{ g}}{1\text{ mg}} \times \frac{1 \times 10^{-3}\text{ kg}}{1\text{ g}} = \mathbf{2.54 \times 10^{-5}}\text{ kg}$

c. $556\text{ mL} \times \frac{1 \times 10^{-3}\ L}{1\ mL} = \mathbf{0.556\,L}$

d. $\frac{10.6\text{ kg}}{1\text{ m}^3} \times \frac{1000\text{ g}}{1\text{ kg}} \times \left(\frac{1 \times 10^{-2}\text{ m}}{1\text{ cm}}\right)^3 = \mathbf{0.0106\,g/cm^3}$

1.39 a. **Strategy:** The solution requires a two-step dimensional analysis because we must first convert pounds to grams and then grams to milligrams.

Setup: The necessary conversion factors as derived from the equalities: 1 g = 1000 mg and 1 lb = 453.6 g.

$$\frac{453.6\text{ g}}{1\text{ lb}} \text{ and } \frac{1\text{ mg}}{1 \times 10^{-3}\text{ g}}$$

Solution:

$$242\text{ lb} \times \frac{453.6\text{ g}}{1\text{ lb}} \times \frac{1\text{ mg}}{1 \times 10^{-3}\text{ g}} = \mathbf{1.10 \times 10^{8}\ mg}$$

b. **Strategy:** We need to convert from cubic centimeters to cubic meters.

Setup: 1 m = 100 cm. When a unit is raised to a power, the corresponding conversion factor must also be raised to that power in order for the units to cancel.

Solution:

$$68.3\ \text{cm}^3 \times \left(\frac{1\ \text{m}}{100\ \text{cm}}\right)^3 = \mathbf{6.83 \times 10^{-5}\ m^3}$$

c. **Strategy:** In Chapter 1 of the text, a conversion is given between liters and cm^3 (1 L = 1000 cm^3). If we can convert m^3 to cm^3, we can then convert to liters. Recall that 1 cm = 1 × 10^{-2} m. We need to set up two conversion factors to convert from m^3 to L. Arrange the appropriate conversion factors so that m^3 and cm^3 cancel, and the unit liters is obtained in your answer.

Setup: The sequence of conversions is $\text{m}^3 \rightarrow \text{cm}^3 \rightarrow$ L. Use the following conversion factors:

$$\left(\frac{1\ \text{cm}}{1 \times 10^{-2}\ \text{m}}\right)^3 \text{ and } \frac{1\ \text{L}}{1000\ \text{cm}^3}$$

Solution:

$$7.2\ \text{m}^3 \times \left(\frac{1\ \text{cm}}{1 \times 10^{-2}\ \text{m}}\right)^3 \times \frac{1\ \text{L}}{1000\ \text{cm}^3} = \mathbf{7.2 \times 10^3\ L}$$

d. **Strategy:** A relationship between pounds and grams is given on the end sheet of the text (1 lb = 453.6 g). This relationship will allow conversion from grams to pounds. If we can convert from micrograms to grams, we can then convert from grams to pounds. Recall that 1 μg = 1 × 10^{-6} g. Arrange the appropriate conversion factors so that micrograms and grams cancel, and the unit pounds is obtained in your answer.

Setup: The sequence of conversions is μg → g → lb. Use the following conversion factors:

$$\frac{1 \times 10^{-6}\ \text{g}}{1\ \mu\text{g}} \text{ and } \frac{1\ \text{lb}}{453.6\ \text{g}}$$

Solution:

$$28.3\ \mu\text{g} \times \frac{1 \times 10^{-6}\ \text{g}}{1\ \mu\text{g}} \times \frac{1\ \text{lb}}{453.6\ \text{g}} = \mathbf{6.24 \times 10^{-8}\ lb}$$

1.40 a. **Strategy:** The given unit is amu and the desired unit is grams. Use a conversion factor to convert amu → g.

Setup: Use the conversion factor:

$$\frac{1.6605378 \times 10^{-24}\text{ g}}{1\text{ amu}}$$

Solution:

$$242\text{ amu} \times \left(\frac{1.6605378 \times 10^{-24}\text{g}}{1\text{ amu}}\right) = \mathbf{4.02 \times 10^{-22}\text{ g}}$$

b. **Strategy:** The given unit is amu and the desired unit is kilograms. Use a conversion factor to convert amu → kg.

Setup: Use the conversion factor:

$$\frac{1.6605378 \times 10^{-27}\text{ kg}}{1\text{ amu}}$$

Solution:

$$87\text{ amu} \times \frac{1.6605378 \times 10^{-27}\text{ kg}}{1\text{ amu}} = \mathbf{1.4 \times 10^{-25}\text{ kg}}$$

c. **Strategy:** The given unit is Å and the desired unit is meters. Use a conversion factor to convert Å → m.

Setup: Use the conversion factor:

$$\frac{1 \times 10^{-10}\text{ m}}{1\text{ Å}}$$

Solution:

$$2.21\text{ Å} \times \frac{1 \times 10^{-10}\text{ m}}{1\text{Å}} = \mathbf{2.21 \times 10^{-10}\text{ m}}$$

d. **Strategy:** The given unit is Å and the desired unit is nanometers. Use conversion factors to conver Å → m and m → nm.

Setup: Use the conversion factors:

$$\frac{1 \times 10^{-10}\text{ m}}{1\text{ Å}} \text{ and } \frac{1\text{ nm}}{1 \times 10^{-9}\text{ m}}$$

Solution:

$$1.73\text{ Å} \times \frac{1 \times 10^{-10}\text{ m}}{1\text{Å}} \times \frac{1\text{ nm}}{1 \times 10^{-9}\text{ m}} = \mathbf{0.173\text{ nm}}$$

1.41 a. **Strategy:** The given unit is grams and the desired unit is amu. Use a conversion factor to convert g → amu.

Setup: Use the conversion factor:

$$\frac{1\ \text{amu}}{1.6605378 \times 10^{-24}\ \text{g}}$$

Solution:

$$1.1 \times 10^{-22}\ \text{g} \times \frac{1\ \text{amu}}{1.6605378 \times 10^{-24}\ \text{g}} = \mathbf{66\ amu}$$

b. **Strategy:** The given unit is kilograms and the desired unit is amu. Use a conversion factor to convert kg → amu.

Setup: Use the conversion factor:

$$\frac{1\ \text{amu}}{1.6605378 \times 10^{-27}\ \text{kg}}$$

Solution:

$$1.08 \times 10^{-29}\ \text{kg} \times \frac{1\ \text{amu}}{1.6605378 \times 10^{-27}\ \text{kg}} = \mathbf{6.50 \times 10^{-3}\ amu}$$

c. **Strategy:** The given unit is meters and the desired unit is Å. Use a conversion factor to convert m → Å.

Setup: Use the conversion factor:

$$\frac{1\ \text{Å}}{1 \times 10^{-10}\ \text{m}}$$

Solution:

$$8.3 \times 10^{-9}\ \text{m} \times \frac{1\text{Å}}{1 \times 10^{-10}\ \text{m}} = \mathbf{83Å}$$

d. **Strategy:** The given unit is picometers and the desired unit is angstroms. Use conversion factors to convert pm → m → Å.

Setup: Use the conversion factors:

$$\frac{1 \times 10^{-12}\ \text{m}}{\text{pm}} \text{ and } \frac{1\text{Å}}{1 \times 10^{-10}\ \text{m}}$$

Solution:

$$132\ \text{pm} \times \frac{1 \times 10^{-12}\ \text{m}}{\text{pm}} \times \frac{1\text{Å}}{1 \times 10^{-10}\ \text{m}} = \mathbf{1.32Å}$$

1.42 $\dfrac{1255\text{ m}}{1\text{ s}} \times \dfrac{1\text{ mi}}{1609\text{ m}} \times \dfrac{3600\text{ s}}{1\text{ h}} = \mathbf{2808\ mi/h}$

1.43 **Strategy:** You should know conversion factors that will allow you to convert between days and hours, hours and minutes, and minutes and seconds. Make sure to arrange the conversion factors so that days, hours, and minutes cancel, leaving units of seconds for the answer.

Setup: The sequence of conversions is d → h → min → s. Use the following conversion factors:

$$\frac{24\text{ h}}{1\text{ d}},\ \frac{60\text{ min}}{1\text{ h}},\ \text{and}\ \frac{60\text{ s}}{1\text{ min}}$$

Solution:

$$365.24\text{ d} \times \frac{24\text{ h}}{1\text{ d}} \times \frac{60\text{ min}}{1\text{ h}} \times \frac{60\text{ s}}{1\text{ min}} = \mathbf{3.1557 \times 10^7\ s}$$

1.44 $$(93 \times 10^6\text{ mi}) \times \frac{1.609\text{ km}}{1\text{ mi}} \times \frac{1000\text{ m}}{1\text{ km}} \times \frac{1\text{ s}}{3.00 \times 10^8\text{ m}} \times \frac{1\text{ min}}{60\text{ s}} = \mathbf{8.3\ min}$$

1.45 a. **Strategy:** The measurement is given in mi/min. We are asked to convert this rate to in/s. Use conversion factors to convert mi → ft → in and to convert min → s.

Setup: Use the conversion factors:

$$\frac{5280\text{ ft}}{1\text{ mi}},\ \frac{12\text{ in}}{1\text{ ft}},\ \text{and}\ \frac{1\text{ min}}{60\text{ s}}$$

Be sure to set the conversion factors up so that the appropriate units cancel.

Solution:

$$\frac{1\text{ mi}}{13\text{ min}} \times \frac{5280\text{ ft}}{1\text{ mi}} \times \frac{12\text{ in}}{1\text{ ft}} \times \frac{1\text{ min}}{60\text{ s}} = \mathbf{81\ in/s}$$

b. **Strategy:** The measurement is given in mi/min. We are asked to convert this rate to m/min. Use a conversion factor convert mi → m.

Setup: Use the conversion factor:

$$\frac{1609\ \text{m}}{1\ \text{mi}}$$

Solution:

$$\frac{1\ \text{mi}}{13\ \text{min}} \times \frac{1609\ \text{m}}{1\ \text{mi}} = \mathbf{1.2 \times 10^2\ m/min}$$

c. **Strategy:** The measurement is given in mi/min. We are asked to convert this rate to km/h. Use conversion factors to convert mi → m → km and convert min → h.

Setup: Use the conversion factors:

$$\frac{1609\ \text{m}}{1\ \text{mi}},\ \frac{1\ \text{km}}{1000\ \text{m}},\ \text{and}\ \frac{60\ \text{min}}{1\ \text{h}}$$

Solution:

$$\frac{1\ \text{mi}}{13\ \text{min}} \times \frac{1609\ \text{m}}{1\ \text{mi}} \times \frac{1\ \text{km}}{1000\ \text{m}} \times \frac{60\ \text{min}}{1\ \text{h}} = \mathbf{7.4\ km/h}$$

1.46

$$6.0\ \text{ft} \times \frac{1\ \text{m}}{3.28\ \text{ft}} = \mathbf{1.8\ m}$$

$$183\ \text{lb} \times \frac{453.6\ \text{g}}{1\ \text{lb}} \times \frac{1\ \text{kg}}{1000\ \text{g}} = \mathbf{83.0\ kg}$$

1.47 **Strategy:** The rate is given in the units mi/h. The desired units are km/h. Use conversion factors to convert mi → m → km.

Setup: Use the conversion factors:

$$\frac{1609\ \text{m}}{1\ \text{mi}}\ \text{and}\ \frac{1\ \text{km}}{1000\ \text{m}}$$

Solution:

$$\frac{20\ \text{mi}}{1\ \text{h}} \times \frac{1609\ \text{m}}{1\ \text{mi}} \times \frac{1\ \text{km}}{1000\ \text{m}} = \mathbf{32\ km/h}$$

1.48 $\frac{62\ \text{m}}{1\ \text{s}} \times \frac{1\ \text{mi}}{1609\ \text{m}} \times \frac{3600\ \text{s}}{1\ \text{h}} = \mathbf{1.4 \times 10^2\ mph}$

1.49 **Strategy:** We seek to convert meters per second to miles per hour.

Setup: The necessary conversion factors are 1000 m = 1 km, 1 km = 0.6215 mi, 60 s = 1 min, and 60 min = 1 hr.

Solution:

$$\frac{377\text{ m}}{\text{s}} \times \frac{1\text{ km}}{1000\text{ m}} \times \frac{60\text{ s}}{1\text{ mm}} \times \frac{60\text{ mm}}{1\text{ h}} \times \frac{0.6215\text{ mi}}{1\text{ km}} = \mathbf{843\ mph}$$

1.50 a. $1.42\text{ km} \times \frac{1\text{ mi}}{1.609\text{ km}} = \mathbf{0.883\ mi}$

b. $32.4\text{ yd} \times \frac{36\text{ in}}{1\text{ yd}} \times \frac{2.54\text{ cm}}{1\text{ in}} = \mathbf{2.960\ cm}$

c. $\frac{3.0 \times 10^{10}\text{ cm}}{1\text{ s}} \times \frac{1\text{ in}}{2.54\text{ cm}} \times \frac{1\text{ ft}}{12\text{ in}} = \mathbf{9.8 \times 10^8\ ft/s}$

1.51 a. **Strategy:** The given unit is nm and the desired unit is m. Use a conversion factor to convert nm → m.

Setup: Use the conversion factor:

$$\frac{1 \times 10^{-9}\text{ m}}{1\text{ nm}}$$

Solution:

$$352\text{ nm} \times \frac{1 \times 10^{-9}\text{ m}}{1\text{ nm}} = \mathbf{3.52 \times 10^{-7}\ m}$$

b. **Strategy:** The given unit is yr and the desired unit is s. Use conversion factors to convert yr → d → h → s.

Setup: Use the conversion factors:

$$\frac{365\text{ d}}{1\text{ yr}}, \frac{24\text{ h}}{1\text{ d}}, \text{and} \frac{3600\text{ s}}{1\text{ h}}$$

Solution:

$$13.8 \times 10^9\text{ yr} \times \frac{365\text{ d}}{1\text{ yr}} \times \frac{24\text{ h}}{1\text{ d}} \times \frac{3600\text{ s}}{1\text{ h}} = \mathbf{4.4 \times 10^{17}\ s}$$

c. **Strategy:** The given unit is cm^3 and the desired unit is m^3. Use a conversion factor to convert $cm^3 \rightarrow m^3$.

Setup: Use the conversion factor:

$$\left(\frac{0.01\ \text{m}}{1\ \text{cm}}\right)^3$$

Solution:

$$93.7\ \text{cm}^3 \times \left(\frac{0.01\ \text{m}}{1\ \text{cm}}\right)^3 = \mathbf{9.37 \times 10^{-5}\ m^3}$$

d. **Strategy:** The given unit is m^3 and the desired unit is L. Use conversion factors to convert $m^3 \rightarrow cm^3 \rightarrow L$.

Setup: Use the conversion factors:

$$\left(\frac{1\ \text{cm}}{1 \times 10^{-2}\ \text{m}}\right)^3 \text{ and } \frac{1\ \text{L}}{1000\ \text{cm}^3}$$

Solution:

$$20.50\ \text{m}^3 \times \left(\frac{1\ \text{cm}}{1 \times 10^{-2}\ \text{m}}\right)^3 \times \frac{1\ \text{L}}{1000\ \text{cm}^3} = \mathbf{2.050 \times 10^4\ L}$$

1.52 $$\text{Density} = \frac{1.85\ \text{g}}{1\ \text{cm}^3} \times \frac{1\ \text{kg}}{1000\ \text{g}} \times \left(\frac{1\ \text{cm}}{0.01\ \text{m}}\right)^3 = \mathbf{1850\,kg/m^3}$$

1.53 **Strategy:** The given unit s g/L and the desired unit is g/cm^3. Use a conversion factor to convert $L \rightarrow cm^3$.

Setup: Use the conversion factor:

$$\frac{1\ \text{L}}{1000\ \text{cm}^3}$$

Solution:

$$\frac{674\ \text{g}}{1\ \text{L}} \times \frac{1\ \text{L}}{1000\ \text{cm}^3} = \mathbf{0.674\,g/cm^3}$$

1.54 **Strategy:** The density of the penny can be calculated using the weight percentages of each element and the given densities of each element.

Setup:

$$\text{Density of penny} = \sum\left(\frac{\text{percent metal}}{100} \times \text{density metal}\right)$$

Solution: **1825**:

The penny was pure copper. Therefore, the density of the penny was that of copper, **8.92 g/cm³**.

1860:

The penny was 88% copper and 12% nickel.

$$\left(\frac{88}{100}\right)\left(8.92\ \text{g/cm}^3\right) + \left(\frac{12}{100}\right)\left(8.91\ \text{g/cm}^3\right) = \mathbf{8.9\ g/cm^3}$$

1965:

The penny was 95% copper and 5% zinc.

$$\left(\frac{95}{100}\right)\left(8.92\ \text{g/cm}^3\right) + \left(\frac{5}{100}\right)\left(7.14\ \text{g/cm}^3\right) = \mathbf{8.8\ g/cm^3}$$

Today:

The penny is 97.5% zinc and 2.5% copper.

$$\left(\frac{97.5}{100}\right)\left(7.14\ \text{g/cm}^3\right) + \left(\frac{2.5}{100}\right)\left(8.92\ \text{g/cm}^3\right) = \mathbf{7.18\ g/cm^3}$$

1.55 Answers may vary.
Samples:

a. Matter – **rock**
b. Substance – **water**
c. Mixture – **seawater**

1.56 Answers may vary.
Samples:

a. Homogeneous mixture – **sugar dissolved in water**
b. Heterogeneous mixture – **sand and iron filings**

1.57 a. **liquid** b. **Gas** c. **mixture** d. **Solid**

1.58 a. The sea is a heterogeneous mixture of seawater and biological matter, but seawater with the biomass filtered out is a **homogeneous mixture**.

b. **pure substance**

c. **pure substance**

d. **homogeneous mixture**

e. **homogeneous mixture**

f. **pure substance**

g. **heterogeneous mixture**

h. **pure substance**

i. **pure substance**

1.59 **A qualitative property of a system does not require explicit measurement. A quantitative property of a system requires measurement and is expressed with a number.**

1.60 **Physical properties can be observed and measured without changing the identity of a substance. For example, the boiling point of water can be determined by heating a container of water and measuring the temperature at which the liquid water turns to steam. The water vapor (steam) is still H_2O, so the identity of the substance has not changed. Liquid water can be recovered by allowing the water vapor to contact a cool surface, on which it condenses to liquid water.**

Chemical properties can be observed only by carrying out a chemical change. During the measurement, the identity of the substance changes. The original substance cannot be recovered by any physical means. For example, when iron is exposed to water and oxygen, it undergoes a chemical change to produce rust. The iron cannot be recovered by any physical means.

1.61 **An extensive property depends on the amount of matter or mass present. An intensive property is independent of the amount of matter present.**

1.62 a. **Extensive**

b. **Extensive**

c. **intensive**

d. **extensive**

1.63 a. **Quantitative**. This statement involves a measurable distance.

b. **Qualitative**. This is a value judgment. There is no numerical scale of measurement for artistic excellence.

c. **Qualitative**. If the numerical values for the densities of ice and water were given, it would be a quantitative statement.

d. **Qualitative**. The statement is a value judgment.

e. **Qualitative**. Even though numbers are involved, they are not the result of measurement.

1.64 a. **Chemical property**. Oxygen gas is consumed in a combustion reaction; its composition and identity are changed.

b. **Chemical property**. The basic ingredients of the antacids undergo a chemical change when they react with stomach acids to reduce acid reflux. The antacids cannot be recovered by a physical process.

c. **Physical property**. The measurement of the boiling point of water does not change its identity or composition.

d. **Physical property**. The measurement of the densities of carbon dioxide and air does not change their composition.

e. **Chemical property**. A chemical change takes place when uranium reacts with fluorine to form a gas. The original substance (uranium) no longer exists, and it cannot be recovered by means of a physical process.

1.65 a. **Physical change.** The material is helium regardless of whether it is located inside or outside the balloon.

b. **Chemical change** in the battery.

c. **Physical change.** The orange juice concentrate can be regenerated by the addition of water.

d. **Chemical change.** Photosynthesis changes water, carbon dioxide, etc., into complex organic matter.

e. **Physical change.** The sugar can be recovered unchanged by evaporation.

1.66 a. **Physical change**. A soda goes flat when the dissolved carbon dioxide gas escapes from the liquid. The carbon dioxide remains unchanged regardless of whether it is dissolved in the soda or it has escaped into the atmosphere. The fizz could be recovered by adding carbon dioxide gas back into the soda.

b. **Chemical change.** A bruise is composed of blood under the skin. The hemoglobin in blood makes it red-blue in color. With time, the hemoglobin breaks down into green-colored biliverdin, yellow-colored bilirubin, and golden-brown hemosiderin. The color of the bruise depends on the amount of each compound present.

c. **Chemical change.** After the leaves are burned during the combustion process, the original substance (leaves, in this case) no longer exists.

d. **Physical change.** When frost forms, water changes its state of matter from a liquid to a solid, but its identity is not changed. Liquid water could be recovered by melting the frost.

e. **Physical change.** Water from the clothes evaporates into water vapor dispersed in the air. Water changes its state from liquid to gas, but its identity is not changed. Liquid water could be recovered by cooling the air.

1.67 a. **Upper ruler:2.5 cm** b. **Lower ruler:2.55 cm**

1.68 Volume of sample: V = 18.45 mL–17.25 mL = 1.20 mL

$$\text{Density: } d = \frac{m}{V} = \frac{10.5\text{ g}}{1.20\text{ mL}} = \mathbf{8.75\ g/mL}$$

1.69 a. **chemical** b. **Chemical** c. **physical** d. **physical** e. **chemical**

1.70

Substance	Qualitative Statement	Quantitative Statement
a. water	**colorless liquid**	**freezes at 0°C**
b. carbon	**black solid (graphite)**	**density = 2.26 g/cm^3**
c. iron	**rusts easily**	**density = 7.86 g/cm^3**
d. hydrogen gas	**colorless gas**	**melts at −259.1°C**
e. sucrose	**tastes sweet**	**at 0°C, 179 g of sucrose dissolves in 100 g of H_2O**
f. salt	**tastes salty**	**melts at 801°C**
g. mercury	**liquid at room temperature**	**boils at 357°C**
h. gold	**a precious metal**	**density = 19.3 g/cm^3**
i. air	**a mixture of gases**	**contains 20% oxygen by volume**

1.71

$$(95.0 \times 10^9\text{ lb sulfuric acid}) \times \frac{1\text{ ton}}{2.0 \times 10^3\text{ lb}} = \mathbf{4.75 \times 10^7\ tons\ of\ sulfuric\ acid}$$

1.72 Volume of rectangular solid $= l \times w \times h$

$$\text{Volume } = (8.53\text{ cm})(2.4\text{ cm})(1.0\text{ cm}) = 20.472\text{ cm}^3$$

$$d = \frac{m}{V} = \frac{52.7064\text{ g}}{20.472\text{ cm}^3} = \mathbf{2.6\ g/cm^3}$$

1.73 a. **Strategy:** Calculate the volume of the sphere using:

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

Then use the density equation, $d = \frac{m}{V}$, to find the mass.

Setup: Solve the density equation for *m* to get *m* = *dV*. Find the volume and substitute it into the equation for *m*.

Solution:

$$V = \left(\frac{4}{3}\right)(3.14159)(10.0\text{ cm})^3 = 4189\text{ cm}^3$$

$$m = dV = \frac{19.3\text{ g}}{1\text{ cm}^3} \times 4189\text{ cm}^3 = \mathbf{8.08 \times 10^4\ g}$$

b. **Strategy:** Compute the volume of the cube using:

$$V = s^3$$

Then, find the mass using the density equation:

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

Setup: Solve the density equation for *m* to get *m* = *dV*. Find the volume and substitute it into the equation for *m.*

Solution: The edge of the cube is *s* = 0.040 mm = 0.0040 cm, and $V = (0.0040\text{ cm})^3 = 6.4 \times 10^{-8}\text{ cm}^3$.

$$m = dV = \frac{21.4\text{ g}}{1\text{ cm}^3} \times \left(6.4 \times 10^{-8}\text{ cm}^3\right) = \mathbf{1.4 \times 10^{-6}\ g}$$

c. **Strategy:** Use the density equation:

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

Setup: Solve the density equation for *m* to get *m* = *dV*.

Solution:

$$50.0\text{ mL} \times \frac{0.798\text{ g}}{1\text{ mL}} = 39.9\text{ g}$$

1.74 You are asked to solve for the inner diameter of the tube. If you can calculate the volume that the mercury occupies, you can calculate the radius of the cylinder, $V_{\text{cylinder}} = \pi r^2 h$ (*r* is the inner radius of the cylinder and *h* is the height of the cylinder). The diameter of the cylinder is 2*r*.

$$\text{Volume of Hg filling cylinder} = \frac{\text{mass of Hg}}{\text{density of Hg}}$$

$$\text{Volume of Hg filling cylinder} = \frac{105.5 \text{ g}}{13.6 \text{ g/cm}^3} = 7.76 \text{ cm}^3$$

Next, solve for the radius of the cylinder.

$$\text{Volume of cylinder} = \pi r^2 h$$

$$r = \sqrt{\frac{\text{volume}}{\pi \times h}}$$

$$r = \sqrt{\frac{7.76 \text{ cm}^3}{\pi \times 12.7 \text{ cm}}} = 0.441 \text{ cm}$$

The diameter of the cylinder equals $2r$.

Diameter of the cylinder = $2r$ = 2(0.441 cm) = **0.882 cm**

1.75 **Strategy:** The difference between the masses of the empty and filled flasks is the mass of the water in the flask. The volume of the water (and the flask) can be found using the density equation.

Setup: Solve the density equation for V:

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

Solution: 87.39 – 56.12 = 31.27 g water

$$V = \frac{m}{d} = \frac{31.27 \text{ g}}{0.9976 \text{ g/cm}^3} = \mathbf{31.35\ cm^3}$$

1.76 $\dfrac{343 \text{ m}}{1 \text{ s}} \times \dfrac{1 \text{ mi}}{1609 \text{ m}} \times \dfrac{3600 \text{ s}}{1 \text{ h}} = \mathbf{767\ mph}$

1.77 **Strategy:** The volume of the piece of platinum is the same as the volume of water it displaces. Once the volume is found, use the density equation to compute the density.

Setup: V = 198.0 – 187.1 = 10.9 mL

Solution:

$$d = \left(\frac{234.0 \text{ g}}{10.5 \text{ mL}}\right) = \mathbf{21.5\ g/mL}$$

The density of a solid is generally reported in g/cm^3 (1 mL = 1 cm^3). Therefore, the density is reported as **21.5 g/cm^3**.

1.78 **Place the ice cube in a beaker containing other liquid, such as an alcohol. The liquid must be *less dense* than the ice in order for the ice to sink. The temperature of the experiment must be maintained at or below 0°C to prevent the ice from melting.**

1.79 **Strategy:** Use the density equation.

Setup:

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

Solution:

$$d = \frac{m}{V} = \frac{2.17 \times 10^3 \text{ g}}{242.2 \text{ cm}^3} = \mathbf{8.96\ g/cm^3}$$

1.80

$$\text{Volume} = \frac{\text{mass}}{\text{density}}$$

$$\textbf{Volume occupied by Li} = \frac{1.20 \times 10^3 \text{ g}}{0.53 \text{ g/cm}^3} = \mathbf{2.3 \times 10^3\ cm^3}$$

1.81 **Strategy:** For the Fahrenheit thermometer, we must convert the possible error of 0.1°F to °C. For each thermometer, use the percent error equation to find the percent error for the measurement.

Setup:

$$0.1°\text{F} \times \frac{5°\text{C}}{9°\text{F}} = 0.056°\text{C}\,.$$

For the Fahrenheit thermometer, we expect $|\text{true value} - \text{experimental value}| = 0.056°\text{C}$.
For the Celsius thermometer, we expect $|\text{true value} - \text{experimental value}| = 0.1°\text{C}$.

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

Solution: For the Fahrenheit thermometer,

$$\textbf{Percent error} = \frac{0.056°\text{C}}{38.9°\text{C}} \times 100\% = \mathbf{0.1\%}$$

For the Celsius thermometer,

$$\textbf{Percent error} = \frac{0.1°\text{C}}{38.9°\text{C}} \times 100\% = \mathbf{0.3\%}$$

1.82 To solve this problem, we need to convert from ft^3 to L. Some tables will have a conversion factor of 28.3 L = 1 ft^3, but we can also calculate it using the dimensional analysis method described in Section 1.6 of the text.

First, convert from ft^3 to L:

$$(5.0 \times 10^7 \text{ ft}^3) \times \left(\frac{12 \text{ in}}{1 \text{ ft}}\right)^3 \times \left(\frac{2.54 \text{ cm}}{1 \text{ in}}\right)^3 \times \frac{1 \text{ mL}}{1 \text{ cm}^3} \times \frac{1 \times 10^{-3} \text{ L}}{1 \text{ mL}} = 1.42 \times 10^9 \text{ L}$$

The mass of vanillin (in g) is:

$$\frac{2.0 \times 10^{-11} \text{ g vanillin}}{1 \text{ L}} \times (1.42 \times 10^9 \text{ L}) = 2.84 \times 10^{-2} \text{ g vanillin}$$

The cost is:

$$(2.84 \times 10^{-2} \text{ g vanillin}) \times \frac{\$112}{50 \text{ g vanillin}} = \mathbf{\$0.064 = 6.4¢}$$

1.83 There are 78.3 + 117.3 = 195.6 Celsius degrees between 0°S and 100°S. We can write this as a conversion factor.

$$\frac{195.6^\circ\text{C}}{100^\circ\text{S}}$$

Set up the equation like a Celsius to Fahrenheit conversion. We need to subtract 117.3°C, because the zero point on the new scale is 117.3°C lower than the zero point on the Celsius scale.

$$? \text{ °C} = \left(\frac{195.6\text{°C}}{100\text{°S}}\right)(? \text{ °S}) - 117.3\text{°C}$$

Solving for ? °S gives:
$$? \text{ °S} = (? \text{ °C} + 117.3\text{°C})\left(\frac{100\text{°S}}{195.6\text{°C}}\right)$$

For 25°C we have:
$$? \text{ °S} = (25\text{°C} + 117.3\text{°C})\left(\frac{100\text{°S}}{195.6\text{°C}}\right) = \mathbf{72.8°S}$$

1.84 **Strategy:** We are asked to determine when °C = °S. Since both °C and °S are used in the conversion factor, we can replace both °C and °S with a common variable, such as °C. Solving the algebraic equation for °C will yield the temperature at which the values are numerically equal.

Setup: Conversion from °C to °S (from Problem 1.83):

$$? \text{°S} = (? \text{°C} + 117.3\text{°C})\left(\frac{100\text{°S}}{195.6\text{°C}}\right)$$

Solution: °C = °S

Replacing °S in the equation with °C yields:

$$? \text{°C} = (? \text{°C} + 117.3\text{°C})\left(\frac{100\text{°C}}{195.6\text{°C}}\right)$$

Combine like terms to yield:

$$\left(1 - \frac{100}{195.6}\right)? \text{°C} = \frac{(117.3)(100)}{195.6}\text{°C}$$

Solving for ? °C gives **123°C**.

1.85 **Strategy:** The key to solving this problem is to realize that all the oxygen needed must come from the 4% difference (20% – 16%) between inhaled and exhaled air. The 240 mL of pure oxygen/min requirement comes from the 4% of inhaled air that is oxygen.

Setup: 240 mL of pure oxygen/min = (0.04)(volume of inhaled air/min)

Solution:

$$\text{Volume of inhaled air/min} = \frac{240 \text{ mL of oxygen/min}}{0.04} = 6000 \text{ mL of inhaled air/min}$$

Since there are 12 breaths per min,

$$\text{Volume of air/breath} = \frac{6000 \text{ mL of inhaled air}}{1 \text{ min}} \times \frac{1 \text{ min}}{12 \text{ breaths}} = \mathbf{5 \times 10^2 \text{ mL/breath}}$$

1.86

a. $$\frac{6000 \text{ mL of inhaled air}}{1 \text{ min}} \times \frac{0.001 \text{ L}}{1 \text{ mL}} \times \frac{60 \text{ min}}{1 \text{ h}} \times \frac{24 \text{ h}}{1 \text{ d}} = \mathbf{8.6 \times 10^3 \text{ L of air/day}}$$

b. $$\frac{8.6 \times 10^3 \text{ L of air}}{1 \text{ d}} \times \frac{2.1 \times 10^{-6} \text{ L CO}}{1 \text{ L of air}} = \mathbf{0.018 \text{ L CO/day}}$$

1.87 **Strategy:** The volume of seawater is given. The strategy is to use the given conversion factors to convert L seawater → g seawater → g NaCl. This result can then be converted to kg NaCl and to tons NaCl. Note that 3.1% NaCl by weight means 100 g seawater = 3.1 g NaCl.

Setup: Use the conversion factors:

$$\frac{1000 \text{ mL seawater}}{1 \text{ L seawater}}, \frac{1.03 \text{ g seawater}}{1 \text{ mL seawater}}, \text{ and } \frac{3.1 \text{ g NaCl}}{100 \text{ g seawater}}$$

Solution:

$$1.5\times10^{21} \text{ L seawater}\times\frac{1000 \text{ mL seawater}}{1 \text{ L seawater}}\times\frac{1.03 \text{ g seawater}}{1 \text{ mL seawater}}\times\frac{3.1 \text{ g NaCl}}{100 \text{ g seawater}} = 4.8\times10^{22} \text{ g NaCl}$$

$$\text{Mass of NaCl (kg)} = 4.8\times10^{22} \text{ g NaCl}\times\frac{1 \text{ kg}}{1000 \text{ g}} = \mathbf{4.8\times10^{19} \text{ kg NaCl}}$$

$$\text{Mass of NaCl (tons)} = 4.8\times10^{22} \text{ g NaCl}\times\frac{1 \text{ lb}}{453.6 \text{ g}}\times\frac{1 \text{ ton}}{2000 \text{ lb}} = \mathbf{5.3\times10^{16} \text{ tons NaCl}}$$

1.88 First, calculate the volume of 1 kg of seawater from the density and the mass. We chose 1 kg of seawater, because the problem gives the amount of Mg in every kilogram of seawater. The density of seawater is given in Problem 1.87.

$$\text{Volume} = \frac{\text{mass}}{\text{density}}$$

$$\text{Volume of 1 kg of seawater} = \frac{1000 \text{ g}}{1.03 \text{ g/mL}} = 970.9 \text{ mL} = 0.9709 \text{ L}$$

In other words, there are 1.3 g of Mg in every 0.9709 L of seawater.

Next, let's convert tons of Mg to grams of Mg.

$$(8.0\times10^{4} \text{ tons Mg})\times\frac{2000 \text{ lb}}{1 \text{ ton}}\times\frac{453.6 \text{ g}}{1 \text{ lb}} = 7.26\times10^{10} \text{ g Mg}$$

Volume of seawater needed to extract 8.0×10^4 tons Mg =

$$(7.26\times10^{10} \text{ g Mg})\times\frac{0.9709 \text{ L seawater}}{1.3 \text{ g Mg}} = \mathbf{5.4\times10^{10} \text{ L of seawater}}$$

1.89 a. **Strategy:** Use the given conversion factor to convert troy oz → g.

Setup: Conversion factor:

$$\frac{31.103 \text{ g Au}}{1 \text{ troy oz Au}}$$

Solution:

$$2.41 \text{ troy oz Au} \times \frac{31.103 \text{ g Au}}{1 \text{ troy oz Au}} = \mathbf{75.0 \text{ g Au}}$$

b. **Strategy:** Use the given conversion factors to convert 1 troy oz → g → lb → oz.

Setup: Conversion factors:

$$\frac{31.103 \text{ g}}{1 \text{ troy oz}}, \frac{1 \text{ lb}}{453.6 \text{ g}}, \text{ and } \frac{16 \text{ oz}}{1 \text{ lb}}$$

Solution:

$$1 \text{ troy oz} \times \frac{31.103 \text{ g}}{1 \text{ troy oz}} \times \frac{1 \text{ lb}}{453.6 \text{ g}} \times \frac{16 \text{ oz}}{1 \text{ lb}} = 1.097 \text{ oz}$$

$$1 \text{ troy oz} = 1.097 \text{ oz}$$

A troy ounce is heavier than an ounce.

1.90

$$\text{Volume} = \text{surface area} \times \text{depth}$$

Recall that 1 L = 1 dm^3. Convert the surface area to units of dm^2 and the depth to units of dm.

$$\text{Surface area} = (1.8 \times 10^8 \text{ km}^2) \times \left(\frac{1000 \text{ m}}{1 \text{ km}}\right)^2 \times \left(\frac{1 \text{ dm}}{0.1 \text{ m}}\right)^2 = 1.8 \times 10^{16} \text{ dm}^2$$

$$\text{Depth} = (3.9 \times 10^3 \text{ m}) \times \frac{1 \text{ dm}}{0.1 \text{ m}} = 3.9 \times 10^4 \text{ dm}$$

$$\text{Volume} = \text{surface area} \times \text{depth} = (1.8 \times 10^{16} \text{ dm}^2)(3.9 \times 10^4 \text{ dm}) = 7.0 \times 10^{20} \text{ dm}^3 = \mathbf{7.0 \times 10^{20} \text{ L}}$$

1.91 a. **Strategy:** Use the percent error equation.

Setup: The percent error of a measurement is given by:

$$\frac{|\text{true value} - \text{experimental value}|}{\text{true value}} \times 100\%$$

Solution:

$$\frac{|0.798 \text{ g/mL} - 0.802 \text{ g/mL}|}{0.798 \text{ g/mL}} \times 100\% = \mathbf{0.5\%}$$

b. **Strategy:** Use the percent error equation.

Setup: The percent error of a measurement is given by:

$$\frac{|\text{true value} - \text{experimental value}|}{\text{true value}} \times 100\%$$

Solution:

$$\frac{|0.864 \text{ g} - 0.837 \text{ g}|}{0.864 \text{ g}} \times 100\% = \mathbf{3.1\%}$$

1.92

$$7.3 \times 10^2 \text{ K} - 273 = \mathbf{4.6 \times 10^{2}\,°C}$$

$$\left((4.6 \times 10^2 \text{ °C}) \times \frac{9\text{°F}}{5\text{°C}}\right) + 32\text{°F} = \mathbf{8.6 \times 10^{2}\,°F}$$

1.93 **Strategy:** Use the percent composition measurement to convert kg ore → g Cu. Note that 34.63% Cu by mass means 100 g ore = 34.63 g Cu.

Setup: Use the conversion factors:

$$\frac{34.63 \text{ g Cu}}{100 \text{ g ore}} \text{ and } \frac{1000 \text{ g}}{1 \text{ kg}}$$

Solution:

$$(5.11 \times 10^3 \text{ kg ore}) \times \frac{34.63 \text{ g Cu}}{100 \text{ g ore}} \times \frac{1000 \text{ g}}{1 \text{ kg}} = \mathbf{1.77 \times 10^6 \text{ g Cu}}$$

1.94

$$8.0 \times 10^4 \text{ tons Au} \times \frac{2000 \text{ lb Au}}{1 \text{ ton Au}} \times \frac{16 \text{ oz Au}}{1 \text{ lb Au}} \times \frac{\$1100}{1 \text{ oz Au}} = \mathbf{\$2.8 \times 10^{12} \text{ or } \$2.8 \text{ trillion}}$$

1.95 **Strategy:** Use the given rates to convert cars → kg CO_2.

Setup: Conversion factors:

$$\frac{5000 \text{ mi}}{1 \text{ car}}, \frac{1 \text{ gal gas}}{20 \text{ mi}}, \text{ and } \frac{9.5 \text{ kg CO}_2}{1 \text{ gal gas}}$$

Solution:

$$(40 \times 10^6 \text{ cars}) \times \frac{5000 \text{ mi}}{1 \text{ car}} \times \frac{1 \text{ gal gas}}{20 \text{ mi}} \times \frac{9.5 \text{ kg CO}_2}{1 \text{ gal gas}} = \mathbf{9.5 \times 10^{10} \text{ kg CO}_2}$$

1.96

$$\text{volume} = \text{area} \times \text{thickness}$$

From the density, we can calculate the volume of the Al foil.

$$\text{volume} = \frac{\text{mass}}{\text{density}} = \frac{3.636 \text{ g}}{2.699 \text{ g/cm}^3} = 1.347 \text{ cm}^3$$

Convert the unit of area from ft^2 to cm^2.

$$1.000 \text{ ft}^2 \times \left(\frac{12 \text{ in}}{1 \text{ ft}}\right)^2 \times \left(\frac{2.54 \text{ cm}}{1 \text{ in}}\right)^2 = 929.0 \text{ cm}^2$$

$$\text{thickness} = \frac{\text{volume}}{\text{area}} = \frac{1.347 \text{ cm}^3}{929.0 \text{ cm}^2} = 1.450 \times 10^{-3} \text{ cm} = \mathbf{1.450 \times 10^{-2} \text{ mm}}$$

1.97 **Strategy:** Use the given rate to convert J → yr.

Setup: Conversion factor:

$$\frac{1 \text{ yr}}{1.8 \times 10^{20} \text{ J}}$$

Solution:

$$(2.0 \times 10^{22} \text{ J}) \times \frac{1 \text{ yr}}{1.8 \times 10^{20} \text{ J}} = \mathbf{1.1 \times 10^2 \text{ yr}}$$

1.98 **Strategy:** The given units are amu and angstrom and the desired units are gram and meter. Use conversion factors to convert amu → g and Å → m.

Setup: Use the conversion factors:

$$\frac{1.6605378 \times 10^{-24}\ \text{g}}{1\ \text{amu}} \text{ and } \frac{1 \times 10^{-10}\ \text{m}}{1\ \text{Å}}$$

Solution:

$$308{,}859\ \text{amu} \times \frac{1.6605378 \times 10^{-24}\ \text{g}}{1\ \text{amu}} = \mathbf{5.12872 \times 10^{-19}\ g}$$

$$22\text{Å} \times \frac{1 \times 10^{-10}\ \text{m}}{1\text{Å}} = \mathbf{2.2 \times 10^{-9}\ m}$$

$$26\text{Å} \times \frac{1 \times 10^{-10}\ \text{m}}{1\text{Å}} = \mathbf{2.6 \times 10^{-9}\ m}$$

22 Å to 26 Å is equal to a range of 2.2 × 10^{-9} m to 2.6 × 10^{-9} m.

1.99 **Strategy:** To calculate the density of the pheromone, you need the mass of the pheromone and the volume that it occupies. The mass is given in the problem.

Setup: volume of a cylinder = area × height = $\pi r^2 \times h$

Converting the radius and height to cm gives:

$$0.50\ \text{mi} \times \frac{1609\ \text{m}}{1\ \text{mi}} \times \frac{1\ \text{cm}}{0.01\ \text{m}} = 8.05 \times 10^4\ \text{cm}$$

$$40\ \text{ft} \times \frac{12\ \text{in}}{1\ \text{ft}} \times \frac{2.54\ \text{cm}}{1\ \text{in}} = 1.22 \times 10^3\ \text{cm}$$

Solution: volume = $\pi(8.05 \times 10^4 \text{ cm})^2 \times (1.22 \times 10^3 \text{ cm}) = 2.48 \times 10^{13} \text{ cm}^3$

Density of gases is usually expressed in g/L. Let's convert the volume to liters.

$$(2.48 \times 10^{13} \text{ cm}^3) \times \frac{1 \text{ mL}}{1 \text{ cm}^3} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 2.48 \times 10^{10} \text{ L}$$

$$\textbf{density} = \frac{\text{mass}}{\text{volume}} = \frac{1.0 \times 10^{-8} \text{ g}}{2.48 \times 10^{10} \text{ L}} = \mathbf{4.0 \times 10^{-19}\ g/L}$$

1.100 First, calculate the mass (in g) of water in the pool. We perform this conversion because we know there is 1 g of chlorine needed per million grams of water.

$$(2.0 \times 10^4 \text{ gallons } H_2O) \times \frac{3.79 \text{ L}}{1 \text{ gallon}} \times \frac{1 \text{ mL}}{0.001 \text{ L}} \times \frac{1 \text{ g}}{1 \text{ mL}} = 7.58 \times 10^7 \text{ g } H_2O$$

Next, let's calculate the mass of chlorine that needs to be added to the pool.

$$(7.58 \times 10^7 \text{ g } H_2O) \times \frac{1 \text{ g chlorine}}{1 \times 10^6 \text{ g } H_2O} = 75.8 \text{ g chlorine}$$

The chlorine solution is only 6.0% chlorine by mass. We can now calculate the volume of chlorine solution that must be added to the pool.

$$75.8 \text{ g chlorine} \times \frac{100\% \text{ solution}}{6.0\% \text{ chlorine}} \times \frac{1 \text{ mL solution}}{1 \text{ g solution}} = \mathbf{1.3 \times 10^3\ mL\ of\ solution}$$

1.101 **Strategy:** We wish to calculate the density and radius of the ball bearing. For both calculations, we need the volume of the ball bearing. The data from the first experiment can be used to calculate the density of the mineral oil. In the second experiment, the density of the mineral oil can then be used to determine what part of the 40.00 mL volume is due to the mineral oil and what part is due to the ball bearing. Once the volume of the ball bearing is determined, we can calculate its density and radius.

Solution: From the first experiment:

$$\text{Mass of oil} = 159.446 \text{ g} - 124.966 \text{ g} = 34.480 \text{ g}$$

$$\text{Density of oil} = \frac{34.480 \text{ g}}{40.00 \text{ mL}} = 0.8620 \text{ g/mL}$$

From the second experiment:

$$\text{Mass of oil} = 50.952 \text{ g} - 18.713 \text{ g} = 32.239 \text{ g}$$

$$\text{Volume of oil} = 32.239 \text{ g} \times \frac{1 \text{ mL}}{0.8620 \text{ g}} = 37.40 \text{ mL}$$

The volume of the ball bearing is obtained by difference.

$$\text{Volume of ball bearing} = 40.00 \text{ mL} - 37.40 \text{ mL} = 2.60 \text{ mL} = 2.60 \text{ cm}^3$$

Now that we have the volume of the ball bearing, we can calculate its density and radius.

$$\text{Density of ball bearing} = \frac{18.713 \text{ g}}{2.60 \text{ cm}^3} = \mathbf{7.20 \text{ g/cm}^3}$$

Using the formula for the volume of a sphere, we can solve for the radius of the ball bearing.

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

$$2.60 \text{ cm}^3 = \frac{4}{3}\pi r^3$$

$$r^3 = 0.621 \text{ cm}^3$$

$$\boldsymbol{r = 0.853 \text{ cm}}$$

1.102 We assume that the thickness of the oil layer is equivalent to the length of one oil molecule. We can calculate the thickness of the oil layer from the volume and surface area.

$$40 \text{ m}^2 \times \left(\frac{1 \text{ cm}}{0.01 \text{ m}}\right)^2 = 4.0 \times 10^5 \text{ cm}^2$$

$$0.10 \text{ mL} = 0.10 \text{ cm}^3$$

$$\text{Volume} = \text{surface area} \times \text{thickness}$$

$$\text{Thickness} = \frac{\text{volume}}{\text{surface area}} = \frac{0.10 \text{ cm}^3}{4.0 \times 10^5 \text{ cm}^2} = 2.5 \times 10^{-7} \text{ cm}$$

Converting to nm:

$$(2.5 \times 10^{-7} \text{ cm}) \times \frac{0.01 \text{ m}}{1 \text{ cm}} \times \frac{1 \text{ nm}}{1 \times 10^{-9} \text{ m}} = \mathbf{2.5 \text{ nm}}$$

1.103 **It would be more difficult to prove that the unknown substance is a pure substance. Most mixtures can be separated into two or more substances.** For example, upon distillation a mixture of salt and water can be separated.

1.104

a. $$\frac{\$1.30}{15.0\ \text{ft}^3} \times \left(\frac{1\ \text{ft}}{12\ \text{in}}\right)^3 \times \left(\frac{1\ \text{in}}{2.54\ \text{cm}}\right)^3 \times \frac{1\ \text{cm}^3}{1\ \text{mL}} \times \frac{1\ \text{mL}}{0.001\ \text{L}} = \mathbf{\$3.06 \times 10^{-3}/L}$$

b. $$2.1\ \text{L water} \times \frac{0.304\ \text{ft}^3\ \text{gas}}{1\ \text{L water}} \times \frac{\$1.30}{15.0\ \text{ft}^3\text{gas}} = \mathbf{\$0.055} = \mathbf{5.5¢}$$

1.105 **Gently heat the liquid to see if any solid remains after the liquid evaporates. Also, collect the vapor and then compare the densities of the condensed liquid with the original liquid. The composition of amixed liquid frequently changes with evaporation along with its density.**

1.106 This problem is similar in concept to a limiting reactant problem. We need sets of coins with three quarters, one nickel, and two dimes. First, we need to find the total number of each type of coin.

$$\text{Number of quarters} = (33.871 \times 10^3\ \text{g}) \times \frac{1\ \text{quarter}}{5.645\ \text{g}} = 6000\ \text{quarters}$$

$$\text{Number of nickels} = (10.432 \times 10^3\ \text{g}) \times \frac{1\ \text{nickel}}{4.967\ \text{g}} = 2100\ \text{nickels}$$

$$\text{Number of dimes} = (7.990 \times 10^3\ \text{g}) \times \frac{1\ \text{dime}}{2.316\ \text{g}} = 3450\ \text{dimes}$$

Next, we need to find which coin limits the number of sets that can be assembled. For each set of coins, we need two dimes for every one nickel.

$$2100\ \text{nickels} \times \frac{2\ \text{dimes}}{1\ \text{nickel}} = 4200\ \text{dimes}$$

We do not have enough dimes.

For each set of coins, we need two dimes for every three quarters.

$$6000\ \text{quarters} \times \frac{2\ \text{dimes}}{3\ \text{quarters}} = 4000\ \text{dimes}$$

Again, we do not have enough dimes, and therefore the number of dimes is our "limiting reactant".

If we need two dimes per set, the number of sets that can be assembled is:

$$3450 \text{ dimes} \times \frac{1 \text{ set}}{2 \text{ dimes}} = 1725 \text{ sets}$$

The mass of each set is:

$$\left(3 \text{ quarters} \times \frac{5.645 \text{ g}}{1 \text{ quarter}}\right) + \left(1 \text{ nickel} \times \frac{4.967 \text{ g}}{1 \text{ nickel}}\right) + \left(2 \text{ dimes} \times \frac{2.316 \text{ g}}{1 \text{ dime}}\right) = 26.534 \text{ g/set}$$

Finally, the total mass of 1725 sets of coins is:

$$1725 \text{ sets} \times \frac{26.534 \text{ g}}{1 \text{ set}} = \mathbf{4.577 \times 10^4 \text{ g}}$$

1.107 **Strategy:** As water freezes, it expands. First, calculate the mass of the water at 20°C. Then, determine the volume that this mass of water would occupy at −5°C.

Solution:

$$\text{Mass of water} = 242 \text{ mL} \times \frac{0.998 \text{ g}}{1 \text{ mL}} = 241.5 \text{ g}$$

$$\text{Volume of ice at } -5°\text{C} = 241.5 \text{ g} \times \frac{1 \text{ mL}}{0.916 \text{ g}} = 264 \text{ mL}$$

The volume occupied by the ice is larger than the volume of the glass bottle. The glass bottle would break.

1.108 We want to calculate the mass of the cylinder, which can be calculated from its volume and density. The volume of a cylinder is $\pi r^2 l$. The density of the alloy can be calculated using the mass percentages of each element and the given densities of each element.

The volume of the cylinder is:

$$V = \pi r^2 l$$

$$V = \pi(6.44 \text{ cm})^2(44.37 \text{ cm})$$

$$V = 5781 \text{ cm}^3$$

The density of the cylinder is:

$$\text{density} = (0.7942)(8.94\ \text{g/cm}^3) + (0.2058)(7.31\ \text{g/cm}^3) = 8.605\ \text{g/cm}^3$$

Now, we can calculate the mass of the cylinder.

$$\text{mass} = \text{density} \times \text{volume}$$

$$\text{mass} = (8.605\ \text{g/cm}^3)(5781\ \text{cm}^3) = \mathbf{4.97 \times 10^4\ g}$$

The calculation assumes that the volumes of the two components are additive. If the volumes are additive, then the density of the alloy is simply the weighted average of the densities of the components.

1.109 a. **Strategy:** Density is the ratio of mass to volume: $d = \dfrac{m}{V}$

Setup: Reading the volumes to two placed past the decimal point (estimating the last digit) gives an initial volume of 10.80 mL and a final volume of 17.65 mL.

Solution: The volume displaced by the solid is 17.65 mL − 10.80 mL = 6.85 mL.

$$\frac{10.211\ \text{g}}{6.85\ \text{mL}} = 1.49\ \text{g/mL or } 1.49\ \text{g/cm}^3$$

1.110 The density of the mixed solution should be based on the percentage of each liquid and its density. Because the solid object is suspended in the mixed solution, it should have the same density as the solution. The density of the mixed solution is:

$$(0.4137)(2.0514\ \text{g/mL}) + (0.5863)(2.6678\ \text{g/mL}) = \mathbf{2.413\ g/mL}$$

As discussed, the density of the object should have the same density as the mixed solution (**2.413 g/mL**).

Yes, this procedure can be used in general to determine the densities of solids. This procedure is called the flotation method. **It is based on the assumptions that the liquids are totally miscible and that the volumes of the liquids are additive.**

1.111 a. **Strategy:** Substitute the temperature values given in the problem into the appropriate equation.

Setup: Conversion from °F to °C:

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

Conversion from °C to K:

$$\text{K} = °\text{C} + 273.15$$

Solution:

$$°\text{C} = (980°\text{F} - 32°\text{F}) \times \frac{5°\text{C}}{9°\text{F}} = \mathbf{527°C}$$

$$°\text{C} = (2240°\text{F} - 32°\text{F}) \times \frac{5°\text{C}}{9°\text{F}} = \mathbf{1227°C}$$

$$\text{K} = 527°\text{C} + 273 = \mathbf{800\ K}$$

$$\text{K} = 1227°\text{C} + 273 = \mathbf{1500\ K}$$

The range of 700°C is equal to a range of 700K.

Note that when there are no digits to the right of the decimal point in the original temperature, we use 273 instead of 273.15

b. **Strategy:** We know the temperature change takes place over the course of 6 h. To determine the rate of temperature change per second, we need to use conversion factors to convert h → min → s. The rate is then equal to the value of the temperature change divided by the number of seconds.

Setup:

$$\text{Rate of temperature change per second} = \frac{\text{change in temperature}}{\text{number of seconds}}$$

Use the conversion factors:

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

Solution: Fahrenheit:

$$2240°F – 980°F = 1260°F$$

$$\frac{1260°F}{6\text{ h}} \times \frac{1\text{ h}}{60\text{ min}} \times \frac{1\text{ min}}{60\text{ s}} = \mathbf{5.833 \times 10^{-2}°F/s}$$

Celsius:

$$1227°C – 527°C = 700°C$$

$$\frac{700°C}{6\text{ h}} \times \frac{1\text{ h}}{60\text{ min}} \times \frac{1\text{ min}}{60\text{ s}} = \mathbf{3.241 \times 10^{-2}°C/s}$$

Kelvin:

$$1500K – 800K = 700\text{ K}$$

$$\frac{700°K}{6\text{ h}} \times \frac{1\text{ hr}}{60\text{ min}} \times \frac{1\text{ min}}{60\text{ s}} = \mathbf{3.241 \times 10^{-2}°K/s}$$

1.112 When CO_2 is released, the mass of the solution will decrease. If we know the starting mass of the solution and the mass of solution after the reaction is complete (given in the problem), we can calculate the mass of CO_2 produced. Then, using the density of CO_2, we can calculate the volume of CO_2 released.

$$\text{Mass of hydrochloric acid} = 40.00\text{ mL} \times \frac{1.140\text{ g}}{1\text{ mL}} = 45.60\text{ g}$$

$$\text{Mass of solution before reaction} = 45.60\text{ g} + 1.328\text{ g} = 46.928\text{ g}$$

We can now calculate the mass of CO_2 by difference.

$$\text{Mass of } CO_2 \text{ released} = 46.928\text{ g} - 46.699\text{ g} = 0.229\text{ g}$$

Finally, we use the density of CO_2 to convert to liters of CO_2 released.

$$\text{Volume of } CO_2 \text{ released} = 0.229\text{ g} \times \frac{1\text{ L}}{1.81\text{ g}} = \mathbf{0.127\text{ L}}$$

1.113 a. **Strategy:** We are asked to determine when °C = °F. Since both °C and °F are used in the conversion factor, we can replace both °C and °F with a common variable, such as °C. Solving the algebraic equation for °C will yield the temperature at which the values are numerically equal.

Setup: Conversion from °C to °F:

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

Solution: °C = °F

Replacing °F in the equation with °C yields:

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

Combine like terms to yield:

$$-\frac{4}{5}°C = 32$$

Solving for °C gives **–40°**.

b. **Strategy:** We are asked to determine when °F = K. If °F = K, we can set the conversion factors equal to one another. Then, we solve for °C which is the variable common to both equations. This value, converted to both °F and K, will yield the value at which both scales are numerically equivalent.

Setup: Conversion from °C to °F:

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

Conversion from °C to K:

$$K = °C + 273.15$$

Solution: If we set °F = K, then:

$$\left(°C \times \frac{9}{5}\right) + 32 = °C + 273.15$$

Solve the equation for °C:

$$1.8\ °C + 32 = °C + 273.15$$

$$0.8\ °C = 241.15$$

$$°C = 301.44$$

Use this value of °C to solve for K or °F:

$$K = 301.44°C + 273.15 = \mathbf{574.59K}$$

$$°F = \left(301.44°C \times \frac{9°F}{5°C}\right) + 32°F = \mathbf{574.59°F}$$

c. **Strategy:** Use the equation which shows the conversion between °C and K.

Setup: Conversion from °C to K:

$$K = °C + 273.15$$

Solution: **No. Since the value of Kelvin is always equal to a value that is 273.15 greater than the Celsius value, there is no temperature at which the values can be numerically equal.**

1.114 **Strategy:** Substitute the temperature value given in the problem into the appropriate equation.

Setup: Conversion from °C to °F:

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

Conversion from Δ°C to Δ°F:

$$\Delta°F = \Delta°C \times \left(\frac{9°F}{5°C}\right)$$

Solution:

$$°F = 58°C \times \left(\frac{9°F}{5°C}\right) + 32°F = \mathbf{136°F}$$

$$\Delta°F = 7.0°C \times \left(\frac{9°F}{5°C}\right) = \mathbf{12°F}$$

1.115 a. **Strategy:** Substitute the temperature values given in the problem into the equation for converting °C to °F.

Setup: Conversion from °F to °C:

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

Solution:

$$°C = (68°F - 32°F) \times \frac{5°C}{9°F} = 20°C$$

$$°C = (77°F - 32°F) \times \frac{5°C}{9°F} = 25°C$$

Storage temperature range: **20°C to 25°C**

b. **Strategy:** Use the density equation,

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

Solution:

$$d = \frac{m}{V} = \frac{5.89 \text{ g}}{5.00 \text{ mL}} = \mathbf{1.18 \text{ g/mL}}$$

c. **Strategy:** The density was determined in g/mL. We are asked to convert this value to g/L and kg/m^3. For g/L, use conversion factors to convert mL → L. For kg/m^3, use conversion factors to convert g/mL → kg/m^3.

Setup: For mL → L, use the conversion factor:

$$\frac{1 \times 10^3 \text{ mL}}{1 \text{ L}}$$

For g/mL→ kg/m^3:

$$1 \text{ g/mL} = 1000 \text{ kg/m}^3$$

Solution:

$$\frac{1.18 \text{ g}}{\text{mL}} \times \frac{1 \times 10^3 \text{ mL}}{\text{L}} = \mathbf{1180 \text{ g/L}}$$

$$1.18 \text{ g/mL} \times \frac{1000 \text{ kg/m}^3}{1 \text{ g/mL}} = \mathbf{1180 \text{ kg/m}^3}$$

d. **Strategy:** The dosage is given in mg per kg of body weight. The given body weight is in pounds so conversion factors should be used to convert lb → g → kg.

Setup: Use the conversion factors:

$$\frac{453.6 \text{ g}}{1 \text{ lb}} \text{ and } \frac{1 \text{ kg}}{1 \times 10^3 \text{ g}}$$

Solution:

$$185 \text{ lb} \times \frac{453.6 \text{ g}}{1 \text{ lb}} \times \frac{1 \text{ kg}}{1 \times 10^3 \text{ g}} = 83.92 \text{ kg}$$

$$\frac{5 \text{ mg}}{\text{kg}} \times 83.92 \text{ kg} = \mathbf{420 \text{ mg}}$$