CHAPTER 2

The Numerical Side of Chemistry

2.1 See solution in textbook.

2.2 Ike is more accurate. Mike's average value is 262, which is higher than the actual value; Ike's average value is 260, which is equal to the actual value. However, Mike is more precise because his values have a spread of 10 $(266 - 256)$ and Ike's have a spread of 36 $(278 - 242)$.

2.3 Jack will be more accurate. If he completely fills the half-quart container twice, the total volume will be very close to 1 quart. However, Jill needs to estimate $1/40$ of the 10-gallon container, which is difficult to do with much accuracy $(1/40$ because 1 gallon $= 4$ quarts$)$.

2.4 See solution in textbook.

2.5 The uncertainty is ± 0.1 gallon because the last digit in the measured volume, 16.0 gallons, is in the tenths column.

2.6 The uncertainty is ± 0.01 V because the tenths value can be read from the dial (each shorter mark on the dial is 0.1 V). Thus the first digit that must be estimated is the one in the hundredths place.

2.7 See solution in textbook.

2.8 You would express the uncertainty ± 0.1 in. in the measured value 600 in. by using a decimal point—600.—to indicate that both zeros are significant.

2.9

	Number of significant figures	Uncertainty
10.0	3	± 0.1
0.004 60	3	± 0.00001
123	3	± 1

2.10 See solution in textbook.

2.11 0.473 (the negative exponent means the number gets smaller).

2.12 47, 325 (the positive exponent means the number gets larger).

2.13 See solution in textbook.

2.14 0.002 35

2.15 6000

2.16 See solution in textbook.

2.17 $4.710\,000\,0 \times 10^{13}$. The fact that the uncertainty is ± 1 million tells you the final significant digit is in the 1-million column, which in this number is the fifth zero from the left.

2.18 $4.710\,000 \times 10^{13}$. The uncertainty of ± 10 million tells you the last significant digit is in the 10-millions column, the fourth zero from the left.

2.19 See solution in textbook.

2.20 44 miles^2. The answer can have only two significant figures because of the 2.0 miles.

2.21 660. hours. The exact 3 has an infinite number of significant figures, meaning the number of significant figures in the answer is determined by the value 220. hours. The decimal point following the zero tells you this number has three significant figures, and that is how many the answer must have.

2.22 See solution in textbook.

2.23 See solution in textbook.

2.24 (a) 6.1×10^2 pounds/in. The answer can have only two significant figures because of the 2.0 in.

(b) 6.11×10^2 or 611 pounds/in. The answer can have only three significant figures because of the 2.00 in.

(c) 86.88 cm because the 4 you multiply by is an exact number, assumed to have an infinite number of significant figures. Thus the product of 21.72×4 should contain the same number of digits as there are in 21.72.

2.25 See solution in textbook.

2.26

$$\begin{array}{rl} & 1555 \text{ cm} \\ + & 0.001 \text{ cm} \\ + & 0.08 \text{ cm} \\ \hline \end{array}$$

1555.801 cm, which rounded off to the correct number of significant figures is 1556 cm.

2.27

$$\begin{array}{r} 142 \text{ cm} \\ -0.48 \text{ cm} \\ \hline \end{array}$$

141.52 cm, which rounded off to the correct number of significant figures is 142 cm.

2.28 See solution in textbook.

2.29 4.736 km. The fact that 1 km is the same as 1000 m means that 4.736 km is the same as $4.736 \times 1000 \text{ m} = 4736 \text{ m}$.

2.30 25 mm. The fact that 1 mm is the same as 0.001 m means that 25 mm is the same as $25 \times 0.001 \text{ m} = 0.025 \text{ m}$.

2.31 See solution in textbook.

2.32 Because 1 mL is 1/1000 of a liter, multiply the given number of liters by 1000 to get 2.5×10^3 milliliters.

2.33 $1 \text{ cm}^3 = 1 \text{ mL}$, which means that $246.7 \text{ cm}^3 = 246.7 \text{ mL}$.

2.34 $\text{K} = {}^\circ\text{C} + 273.15$; therefore ${}^\circ\text{C} - \text{K} = 273.15$; $263.5 \text{ K} - 273.15 = -9.7\ {}^\circ\text{C}$.

$${}^\circ\text{F} = 32 + \frac{9}{5}{}^\circ\text{C} = 15\ {}^\circ\text{F}$$

2.35 See solution in textbook.

2.36 The volume of the cube is 10.0 mm $\times$ 10.0 mm $\times$ 10.0 mm $= 1.00 \times 10^3$ mm^3. Because the problem asks for grams per milliliter, you must convert this volume to milliliters. The easiest way to do this is to first change mm^3 to cm^3. Note that 10.0 mm $=$ 1.00 cm; thus: $(10.0\text{ mm})^3 = (1.00\text{ cm})^3$ or $1.00 \times 10^3\text{ mm}^3 = 1.00\text{ cm}^3$.

The density of the cube is therefore:

$4.70\text{ g}/1.00\text{ cm}^3 = 4.70\text{ g/cm}^3$.

Because $1.00\text{ cm}^3 = 1.00$ mL, the density is 4.70 g/mL.

2.37 $\dfrac{500.0\text{ g}}{150.5\text{ mL}} = 3.322\text{ g/mL}$

2.38 See solution in textbook.

2.39 $\dfrac{1\text{ day}}{24\text{ h}} \quad \dfrac{24\text{ h}}{1\text{ day}}$

2.40 $50.0\ \cancel{\text{miles}} \times \dfrac{1\text{ h}}{600.0\ \cancel{\text{miles}}} = 0.0833\text{ h}$

2.41 $\dfrac{600.0\text{ miles}}{1\ \cancel{\text{h}}} \times 50.0\ \cancel{\text{h}} = 3.00 \times 10^4\text{ miles}$

2.42 See solution in textbook.

2.43 $500.0\ \cancel{\text{L}} \times \dfrac{1000\ \cancel{\text{mL}}}{1\ \cancel{\text{L}}} \times \dfrac{0.001\ 30\text{ g}}{1\ \cancel{\text{mL}}} = 650.\text{ g} = 6.50 \times 10^2\text{ g}$

$650.\ \cancel{\text{g}} \times \dfrac{1\text{ kg}}{1000\ \cancel{\text{g}}} = 0.650\text{ kg}$

2.44 $1.50\ \cancel{\text{lb}} \times \dfrac{453.6\ \cancel{\text{g}}}{1\ \cancel{\text{lb}}} \times \dfrac{1\text{ mL}}{11.4\ \cancel{\text{g}}} = 59.7\text{ mL}$

2.45 Conversion factors: $\dfrac{6\text{ cups flour}}{1\text{ cake}} \qquad \dfrac{1\text{ cup flour}}{120.0\text{ g flour}}$

$6955\ \cancel{\text{g flour}} \times \dfrac{1\ \cancel{\text{cup flour}}}{120.0\ \cancel{\text{g flour}}} \times \dfrac{1\text{ cake}}{6\ \cancel{\text{cups flour}}} = 9.660\text{ cakes}$

You can bake nine cakes (it’s not possible to bake a partial cake).

2.46 Your time conversion is easy enough—hours to minutes—but going from meters squared to feet squared knowing only the conversion factors given in the chapter means several multiplications plus squaring the factors:

$$250.0\ \frac{\cancel{\text{m}^2}}{\cancel{\text{h}}} \times \frac{1\ \cancel{\text{h}}}{60\text{ min}} \times \left(\frac{100\ \cancel{\text{cm}}}{1\ \cancel{\text{m}}}\right)^2 \times \left(\frac{1\ \cancel{\text{in.}}}{2.54\ \cancel{\text{cm}}}\right)^2 \times \left(\frac{1\text{ ft}}{12\ \cancel{\text{in.}}}\right)^2 = 44.85\text{ ft}^2/\text{min}$$

The answer has four significant digits because 2.54 cm in the centimeter–inch conversion factor is an exact number.

2.47 See solution in textbook.

2.48 See solution in textbook.

2.49 See solution in textbook.

2.50 See solution in textbook.

2.51 See solution in textbook.

2.52 Convert volume in milliliters to mass in grams:

$$50.0\ \cancel{\text{mL}} \times \frac{0.785\ \text{g}}{\cancel{\text{mL}}} = 39.3\ \text{g ethanol}$$

The temperature change is 60.0 °C − 22.0 °C = 38.0 °C. Now get the specific heat of ethanol from Table 2.5 of the textbook and use the heat equation:

$$2.43\frac{\cancel{\text{J}}}{\text{g}\ \cancel{°\text{C}}} \times 39.3\ \text{g} \times 38.0\ \cancel{°\text{C}} \times \frac{1\ \text{kJ}}{1000\ \cancel{\text{J}}} = 3.63\ \text{kJ}$$

2.53 $3.63\ \cancel{\text{kJ}} \times \dfrac{1000\ \text{J}}{1\ \cancel{\text{kJ}}} = 3630\ \text{J}$

$$3.63\ \cancel{\text{kJ}} \times \frac{1\ \text{Cal}}{4.184\ \cancel{\text{kJ}}} = 0.868\ \text{Cal}$$

$$3.63\ \cancel{\text{kJ}} \times \frac{1\ \text{cal}}{4.184\ \cancel{\text{J}}} = \frac{1000\ \cancel{\text{J}}}{1\ \cancel{\text{kJ}}} = 868\ \text{cal}$$

2.54 The temperature change is 35 °C − 22.0 °C = 13.5 °C. Thus

$$1.000\ \frac{\text{cal}}{\cancel{\text{g water}\ °\text{C}}} \times 1.00\ \cancel{\text{kg water}} \times \frac{1000\ \cancel{\text{g water}}}{1\ \cancel{\text{kg water}}} \times 13.5\ \cancel{°\text{C}} = 1.35 \times 10^4\ \text{cal}$$

This is the calorie count for only 0.1000 g of the candy. Because the problem asks for big-C Calories per gram of candy, you have one more step:

$$\frac{1.35 \times 10^4\ \cancel{\text{cal}}}{0.100\ \text{g candy}} \times \frac{1\ \text{Cal}}{1000\ \cancel{\text{cal}}} = 1.35 \times 10^2\ \text{Cal/g candy}$$

2.55 The 3 in “3 ft in a yard” is an exact number and therefore is really 3.0000… , with an unlimited number of significant figures. The 3 in “a certain piece of wood is 3 ft long” comes from a measurement and therefore has some uncertainty associated with it.

2.56 Jack’s answer will be an exact number as there are exactly100 pennies in 1 dollar and one cannot have a partial penny coin. Jill’s answer will be a result of a measurement using such devices as measuring cylinders of various volumes.

2.57 You should choose the accurate result because a precise value that is not accurate is useless. An average of accurate results that were not precise usually gets you closer to the true value. On the other hand, an average of inaccurate but precise results may be far off the true value. 1 in. above the average height measured accurately will be safer than 1 in. above the average height measured precisely, but inaccurately.

2.58 The measurements are precise but inaccurate. The average of the three measurements is 2.5 miles, far from the true value of 1.8 miles, and hence, the measurement is inaccurate. On the other hand, the range between the largest and smallest measurements is only 0.1 mile, which is only a small fraction of the true distance, meaning that the measurements were performed precisely.

2.59 The person with the tape measure. He or she needs to make only one measurement, but the person with the ruler has to make at least 200 measurements and add them to get the length. There would be uncertainty associated with each measurement, resulting in a significant loss of accuracy in the result.

2.60 $\pm 1/16$ in. The uncertain digit is the one that is estimated as lying somewhere between the markings. The ruler is marked in eighths, and therefore the estimating is done in the sixteenths place.

2.61 The uncertainty lies in the last digit written in the number. We often assume an uncertainty of ± 1 in the position of the uncertain digit. Some typical examples:

15.2 cm 15.2 ± 0.1 cm (if measured using a ruler marked with centimeters only)

1534 cm^3 1534 ± 1 cm^3 (if measured using a cylinder marked every 10 cm^3)

0.00987 g 0.00987 ± 0.00001 g (if measured using a so-called analytical balance)

2.62 No, it is not possible because no measuring tool has an infinite number of markings. The last digit written in a reported measured value is always an estimate between the markings.

2.63 (a) 12.60 ± 0.01 cm

(b) 12.6 ± 0.1 cm

(c) $0.000\ 000\ 03 \pm 0.000\ 000\ 01$ inch

(d) 125 ± 1 foot

2.64 (a) Four. (b) Three. (c) One. (d) Three.

2.65 Replacing the uncertain digit, 5, by 1 gives an uncertainty of 0.1 million years (or 100,000 years).

2.66 (a) No trailing zeros. (b) No trailing zeros. (c) No trailing zeros. (d) $0.01\underline{0}$

2.67 (a) $12.2\underline{0}2$ (b) No significant zeros. (c) $2\underline{0}5$ (d) $0.01\underline{0}$

2.68 (a) 12.202 ± 001 km (b) 0.01 ± 0.01 mL (c) 205 ± 1 °C (d) 0.010 ± 0.001 g

2.69 It is not clear whether 30 has one or two significant figures because the zero may or may not be significant. Adding the decimal point at the end of the number indicates that the trailing zero is significant, meaning 30. has two significant digits.

2.70 The measurement 2200 ft can be interpreted as having four, three, or two significant digits. Without more information, you cannot tell.

2.71 (a) 56.0 kg (three significant figures).

(b) 0.000 25 m (two significant figures).

(c) 5,600,000 miles (four significant figures, but you cannot tell that by looking at this standard notation).

(d) 2 ft (one significant figure).

2.72 (a) 56.0 ± 0.1 kg

(b) 0.00025 ± 0.00001 m

(c) $5{,}600{,}000 \pm 1000$ miles

(d) 2 ± 1 ft

2.73 (a) 3×10^1 ft

(b) 3.0×10^1 ft

(c) 3.00×10^1 ft

2.74 2.2×10^3 ft. The uncertainty of 100 ft in scientific notation is $\pm 0.1 \times 10^3$.

2.75 (a) 2.26×10^2 (b) 2.260×10^2 (c) 5.0×10^{-10} (d) 3×10^{-1}
(e) 3.0×10^{-1} (f) 9.00×10^8 (g) 9.000006×10^8

2.76 (a) One significant figure. Uncertainty is ± 0.001 kg.
(b) Two significant figures. Uncertainty is $\pm 0.000\,01$ m.
(c) Three significant figures. Uncertainty is ± 1 L.
(d) Four significant figures. Uncertainty is $\pm 0.000\,001$ m.
(e) Two significant figures. Uncertainty is $\pm 100{,}000$ km.

2.77 102 inches because the least certain measured value, either 100. or 2, has its uncertain digit in the ones position, which means the answer has its uncertain digit in the ones position. The uncertainty is ± 1 in.

2.78 (a) 4.60 cm (b) 4 m^2 (c) 1.001×10^4 J
(d) 1×10^3. Because 0.1 has only one significant figure, your answer can have only one significant figure. Thus, even though your calculator displays 1240, all you are allowed to report is 1×10^3.

2.79 The answer of 3.873143939 miles has too many reported digits. The result of the division of 20,450.2 ft by 5280 ft per mile should be reported as 3.87314 miles (six significant figures because 20,450.2 has six significant digits and 5280 is an exact number).

2.80 (a) 2.55×10^5 km. The $33{,}\underline{3}00$ has its uncertain digit in the hundreds position; the $22\underline{2}{,}000$ has its uncertain digit in the thousands position and so is the less certain value. Therefore the answer must have its uncertain digit in the thousands position: $25\underline{5}{,}300$ becomes 2.55×10^5.
(b) 1.000×10^{18} J. Your display was 1exp18, but both values in this division have four significant figures, meaning the answer should also have four.
(c) 2.11×10^2 m. The uncertain digit is in the ones position in $23\underline{4}$ and in the tenths position in $23.\underline{4}$. The subtraction rule tells you the answer must therefore be uncertain in the ones position: $21\underline{0}.6$ becomes 2.11×10^2.
(d) 4.00×10^4 L. The uncertain digit is in the hundreds position in $4.0\underline{0} \times 10^4 = 40{,}\underline{0}00$ and in the thousandths position in $6.00 \times 10^{-1} = 0.60\underline{0}$. $6.0\underline{0} \times 10^{-1} = 0.60\underline{0}$. The answer must therefore be uncertain in the hundreds position: $40{,}\underline{0}00.600$ becomes 4.00×10^4.

2.81 Length, meter; volume, cubic meter.

2.82 Liter (L) and milliliter (mL). These units are used more often than the cubic meter because they are more commonly encountered in everyday situations and in the laboratory.

2.83 It is always correct to use cm^3 instead of mL. The two units are exactly equivalent.

2.84 To eliminate the confusion caused by having different sets of nonuniform measuring scales.

2.85 (a) 2.31×10^9 m (b) 5.00×10^{-6} m (c) 1.004×10^0 m (d) 5.00×10^{-12} m (e) 2.5×10^2 m

2.86 A Celsius degree is larger. There are only 100 Celsius degrees between the freezing point and boiling point of water. However, there are 180 Fahrenheit degrees in this same temperature range. Therefore a Fahrenheit degree is only $\frac{5}{9}$ the size of a Celsius degree $\left(\frac{100}{180} = \frac{5}{9}\right)$.

2.87 The Celsius and Fahrenheit scales can have negative temperature values. The Kelvin scale cannot because the zero point on the Kelvin scale is absolute zero. There is no colder temperature possible than absolute zero, 0 K.

2.88 (a) $\left(22.5\ ^\circ\mathrm{C} \times \frac{9}{5}\right) + 32 = 72.5\ ^\circ\mathrm{F}$; $22.5\ ^\circ\mathrm{C} + 273.15 = 295.65\ \mathrm{K} = 295.6\ \mathrm{K}$

(b) $(-3.0\ ^\circ\mathrm{F} - 32)\left(\frac{5}{9}\right) = -19.4\ ^\circ\mathrm{C}$; $-19.4\ ^\circ\mathrm{C} + 273.15 = 253.75\ \mathrm{K} = 253.8\ \mathrm{K}$

(c) $100.0\ ^\circ\mathrm{C} + 273.15 = 373.15\ ^\circ\mathrm{C} = 373.2\ \mathrm{K}$; $\left(100.0\ ^\circ\mathrm{C} \times \frac{9}{5}\right) + 32 = 212.0\ ^\circ\mathrm{F}$

(d) $\left(65.1\ ^\circ\mathrm{C} \times \frac{9}{5}\right) + 32 = 149\ ^\circ\mathrm{F}$; $65.1\ ^\circ\mathrm{C} + 273.15 = 338.15\ \mathrm{K} = 338.2\ \mathrm{K}$

2.89 32 °F; 0 °C; 273 K

2.90 −273.15 °C, 0.00 K.

2.91 (a) In the left cylinder, each shorter mark is 0.1 mL, which means the uncertain digit in a volume measurement must be in the hundredths position. The uncertainty is thus ± 0.01 mL. In the right cylinder, each shorter mark is 10 mL, which means the uncertain digit in a volume measurement is in the ones position and the uncertainty is ± 1 mL.

(b) The left cylinder contains 1.18 ± 0.01 mL. The right cylinder contains 98 ± 1 mL. Adding the two numbers yields $98\ \mathrm{mL} + 1.18\ \mathrm{mL} = 99.18\ \mathrm{mL}$, which must be reported as 99 mL because the 98 value restricts your answer to being uncertain in the ones position. The uncertainty in this value is ± 1 mL.

2.92 $V = 2.0\ \mathrm{cm} \times 2.0\ \mathrm{cm} \times 2.0\ \mathrm{cm} = 8.0\ \mathrm{cm}^3$.

2.93 The student who reports 1.5 cm used the ruler incorrectly. The ruler is marked in millimeters, which is tenths of centimeters. The uncertainty therefore lies in the hundredths place, and the measurement should be reported to the hundredths place—1.50 cm.

2.94 The radius is $\frac{2.55\ \mathrm{cm}}{2} = 1.275\ \mathrm{cm} = 1.28\ \mathrm{cm}$.

2.95 Density is the amount of mass in a given volume of a material. It is called a derived unit because it is a combination of one SI base unit, mass, and one SI derived unit, volume.

2.96 From Table 2.4, you know that the density of water at 25 °C is 0.997 g/mL. Therefore,

$$1000.0\ \cancel{\mathrm{mL}} \times 0.997 \frac{\mathrm{g}}{\cancel{\mathrm{mL}}} = 997\ \mathrm{g}$$

2.97 From Table 2.4, you know that the density of mercury at 25 °C is 13.6 g/mL. Therefore,

$$2.0\ \cancel{\mathrm{L}} \times \frac{1000\ \cancel{\mathrm{mL}}}{1\ \cancel{\mathrm{L}}} \times 13.6\ \frac{\mathrm{g}}{\cancel{\mathrm{mL}}} = 27{,}200\ \mathrm{g} = 2.7 \times 10^4\ \mathrm{g}$$

2.98 The volume of the stick is $10.0\ \mathrm{cm} \times 10.0\ \mathrm{cm} \times 10.0\ \mathrm{cm} = 1.00 \times 10^3\ \mathrm{cm}^3 = 1.00 \times 10^3\ \mathrm{mL}$. Therefore:

$$1.00 \times 10^3\ \cancel{\mathrm{mL}} \times \frac{0.9\ \mathrm{g}}{\cancel{\mathrm{mL}}} = 9 \times 10^2\ \mathrm{g}$$

2.99 First determine your own mass using a bathroom scale. One possible way would be then to enter the tub and fill the tub all the way to the top with water, making sure you are completely immersed. Next, carefully come out of the tub, sponging any residual water off yourself back into the tub. Using a measuring cup, refill the tub to the previous level. Keep track of the added volume. Your own volume will be equal to the volume of water that had to be replaced. Calculate density by dividing your mass by the volume of water that had to be replaced.

2.100 The two students measure the same density, 19.3 g/mL. The student who works with the 200-g bar finds that it occupies twice the volume of the 100-g bar. Because density is an intensive property, its value does not depend on the size of the sample.

2.101 Place a mixture of gold and fool's gold in a container filled with liquid mercury, which has a density of 13.6 g/mL. Fool's gold, with a density of 5.02 g/mL, is less dense than the mercury and therefore floats. Gold, with a density of 19.3 g/mL, is denser than the mercury and therefore sinks.

2.102 $0.850\ \cancel{\text{weeks}} \times \dfrac{7\ \cancel{\text{day}}}{1\ \cancel{\text{week}}} \times \dfrac{24\ \cancel{\text{h}}}{1\ \cancel{\text{day}}} \times \dfrac{60\ \cancel{\text{min}}}{1\ \cancel{\text{h}}} \times \dfrac{60\ \text{s}}{1\ \cancel{\text{min}}} = 5.14 \times 10^5\ \text{s}$

2.103 $100.0\ \text{miles} \times \dfrac{1\ \text{h}}{45.0\ \cancel{\text{miles}}} \times \dfrac{60\ \text{min}}{1\ \text{h}} = 133\ \text{min}$

2.104 $\dfrac{25.50\ \text{dollars}}{\cancel{\text{h}}} \times \dfrac{1\ \cancel{\text{h}}}{60\ \cancel{\text{min}}} \times \dfrac{1\ \cancel{\text{min}}}{60\ \text{s}} = 7.083 \times 10^{-3}\ \text{dollars/s}$

2.105 $100.0\ \cancel{\text{glonkins}} \times \dfrac{0.911\ \cancel{\text{ounce}}}{1\ \cancel{\text{glonkin}}} \times \dfrac{28.35\ \cancel{\text{g}}}{1\ \cancel{\text{ounce}}} \times \dfrac{1\ \cancel{\text{mL}}}{19.3\ \cancel{\text{g}}} \times \dfrac{1\ \text{L}}{1000\ \cancel{\text{mL}}} = 0.134\ \text{L}$

2.106 $1.000 \times 10^3\ \cancel{\text{cm}^3} \times \dfrac{1\ \cancel{\text{mL}}}{1\ \cancel{\text{cm}^3}} \times \dfrac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \dfrac{0.264\ \text{gallon}}{1\ \cancel{\text{L}}} = 0.264\ \text{gallon}$

2.107 Note that since $1.00\ \text{cm} = 10.0\ \text{mm}$, $9.56 \times 10^2\ \text{mm} = 95.6\ \text{cm}$. Therefore:

$\text{Volume} = 10.2\ \text{cm} \times 43.7\ \text{cm} \times 95.6\ \text{cm} = 4.26 \times 10^4\ \text{cm}^3 = 4.26 \times 10^4\ \text{mL}.$

$4.26 \times 10^4\ \cancel{\text{mL}} \times \dfrac{1\ \text{L}}{1000\ \cancel{\text{mL}}} = 42.6\ \text{L}$

2.108 The mass in grams is

$2.43 \times 10^2\ \cancel{\text{kg}} \times \dfrac{1000\ \text{g}}{1\ \cancel{\text{kg}}} = 2.43 \times 10^5\text{g}$

The volume is 4.26×10^4 mL (from Problem 2.107)

$\text{Density} = \dfrac{\text{Mass}}{\text{Volume}} = \dfrac{2.43 \times 10^5\ \text{gL}}{4.26 \times 10^4\ \text{mL}} = 5.70\ \text{g/mL}$

2.109 (a) The length of the edge $= 100.0\ \text{cm} + 1.40\ \text{cm} = 101.4\ \text{cm}$. You must report the answer to the tenths place because a sum cannot be more certain than the least certain measurement, which in this case is the 100.0 cm.

(b) $\text{Volume} = (101.4\ \text{cm})^3 = 1.043 \times 10^6\ \text{cm}^3 = 1.043 \times 10^6\ \text{mL}$

(c) $\text{Density} = \dfrac{111\ \cancel{\text{kg}}}{1.043 \times 10^6\ \text{ml}} \times \dfrac{1000\ \text{g}}{1\ \cancel{\text{kg}}} = 0.106\ \text{g/mL}$

2.110 The volume in cubic inches is $6.00\ \text{in.} \times 7.00\ \text{in.} \times 8.00\ \text{in.} = 336\ \text{in}^3$. Because the given conversion factor is for inches, you must cube it:

$336\ \cancel{\text{inches}^3} \times \left(\dfrac{2.54\ \cancel{\text{cm}}}{1\ \cancel{\text{inch}}}\right)^3 \times \dfrac{1\ \cancel{\text{mL}}}{1\ \cancel{\text{cm}^3}} \times \dfrac{1\ \text{L}}{1000\ \cancel{\text{mL}}} = 5.51\ \text{L}$

2.111 You must convert both units of the given speed, and that means many conversion factors. Just take things one step at a time. Start with the numerator, meters to miles; then continue with the denominators, seconds to hours:

$$80.0\,\frac{\cancel{\text{m}}}{\cancel{\text{s}}} \times \frac{3.28\,\cancel{\text{ft}}}{1\,\cancel{\text{m}}} \times \frac{1\text{ mile}}{5280\,\cancel{\text{ft}}} \times \frac{60\,\cancel{\text{s}}}{1\,\cancel{\text{min}}} \times \frac{60\,\cancel{\text{min}}}{1\text{ h}} = 179\text{ miles/h}$$

2.112 In an equation, the two sides are equal to each other and must remain equal in order not to change the meaning of the equation. For the sides to remain equal, whatever is done to one side must also be done to the other. In this case, both sides must be multiplied by the same amount.

2.113 To solve for x means to get x alone on one side of the equals sign—in other words, to *isolate* x. For $y = z/x$, a good first step is to get x out of the denominator and onto the left side, accomplished by multiplying both sides by x:

$$x \times y = \frac{z}{\cancel{x}} \times \cancel{x} \Rightarrow xy = z$$

Dividing both sides by y isolates x:

$$\frac{x\cancel{y}}{\cancel{y}} = \frac{z}{y} \Rightarrow x = \frac{z}{y}$$

(b) First multiply both sides of the equation by x to get it out of the denominator and onto the left side:

$$x \times y = \frac{z}{2\cancel{x}} \times \cancel{x} \Rightarrow xy = \frac{z}{2}$$

Dividing both sides by y isolates x:

$$\frac{x\cancel{y}}{\cancel{y}} = \frac{z}{2y} \Rightarrow x = \frac{z}{2y}$$

2.114 Adding x to both sides gives $y + x = z$. Then subtracting y from both sides gives the value of x:

$$y + x = z - \cancel{x} + \cancel{x} \Rightarrow y + x = z$$

$$\cancel{y} + x - \cancel{y} = z - y \Rightarrow x = z - y$$

2.115 One idea would be to place all x-containing terms on one side of the equation and the plain numbers on the other side of the equation. One can do this by subtracting $3x$ and adding 6 to both sides of the equation. Dividing both sides of the equation by 2 isolates x:

$$5x - 3x - \cancel{6} + \cancel{6} = \cancel{3}x - \cancel{3}x - 8 + 6 \Rightarrow 2x = -2$$

$$\frac{\cancel{2}x}{\cancel{2}} = \frac{-\cancel{2}}{\cancel{2}} \Rightarrow x = -1$$

2.116 Using algebraic manipulation means solving the density equation for mass:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

$$\text{Volume} \times \text{Density} = \cancel{\text{Volume}} \times \frac{\text{Mass}}{\cancel{\text{Volume}}}$$

$$\text{Volume} \times \text{Density} = \text{Mass}$$

Substituting in the given values gives

$$50.00\text{ mL} \times 1.15\text{ g/mL} = 57.5\text{ g}$$

2.117 With unit analysis, start with the information given and multiply by the appropriate conversion factor:

$$\frac{1.15\ \text{g}}{\cancel{\text{mL}}} \times 50.00\ \cancel{\text{mL}} = 57.5\text{g}$$

The answer is the same as in Problem 2.116.

2.118 Energy is the capacity for doing work or transferring heat.

2.119 1 cal is the amount of heat energy necessary to warm 1 g of water from 25 °C to 26 °C.

2.120 (a) $4.50\ \cancel{\text{Cal}} \times \dfrac{1000\ \text{cal}}{1\ \cancel{\text{Cal}}} = 4500\ \text{cal} = 4.50 \times 10^3\ \text{cal}$

(b) $600.0\ \cancel{\text{Cal}} \times \dfrac{4.184\ \text{kJ}}{1\ \cancel{\text{Cal}}} = 2510.\ \text{kJ}$

(c) $1.000\ \cancel{\text{J}} \times \dfrac{1\ \text{Cal}}{4.184\ \cancel{\text{J}}} = 0.2390\ \text{Cal}$

(d) $50.0\ \cancel{\text{Cal}} \times \dfrac{4.184\ \text{kJ}}{1\ \cancel{\text{Cal}}} \times \dfrac{1000\ \text{J}}{1\ \cancel{\text{kJ}}} = 2.09 \times 10^5\ \text{J}$

2.121 The specific heat for any substance is the amount of heat energy necessary to increase the temperature of 1 g of the substance by 1 °C.

2.122 The specific heats are 0.901 J/g · °C for aluminum and 0.449 J/g · °C for iron.

$$\textit{Heat}(\text{iron}) = \frac{0.449\ \text{J}}{\text{g} \cdot {}^\circ\text{C}} \times 100.0\ \text{g} \times 75.0\ {}^\circ\text{C} = 3.37 \times 10^3\ \text{J}$$

$$\textit{Heat}(\text{aluminum}) = \frac{0.901\ \text{J}}{\text{g} \cdot {}^\circ\text{C}} \times 100.0\ \text{g} \times 75.0\ {}^\circ\text{C} = 6.76 \times 10^3\ \text{J}$$

The block of aluminum needs $6.76 \times 10^3\ \text{J} - 3.37 \times 10^3\ \text{J} = 3.39 \times 10^3\ \text{J} = 3.39\ \text{kJ}$ more heat than the block of iron.

2.123 Use specific heat of water of 4.184 J/g · °C from Table 2.5:

$$\underbrace{2.00\ \cancel{\text{L}} \times \frac{1000\ \cancel{\text{mL}}}{1\ \cancel{\text{L}}} \times \frac{1.00\ \cancel{\text{g}}}{\cancel{\text{mL}}}}_{\text{Water mass}} \times \underbrace{\frac{4.184\ \text{J}}{\cancel{\text{g}} \cdot \cancel{{}^\circ\text{C}}}}_{\substack{\text{Specific}\\ \text{heat of water}}} \times \underbrace{18.0\ \cancel{{}^\circ\text{C}}}_{\substack{\text{Temperature}\\ \text{increase}}} = 1.51 \times 10^5\ \text{J}$$

$$1.51 \times 10^5\ \cancel{\text{J}} \times \frac{1\ \text{kJ}}{1000\ \cancel{\text{J}}} = 151 \times 10^2\ \text{kJ}$$

$$1.51 \times 10^2\ \text{kJ} \times \frac{1\ \text{Cal}}{4.184\ \text{kJ}} = 36.0\ \text{Cal}$$

Alternatively, you could use the specific heat of water of 1.000 cal/g · °C from Table 2.5:

$$\underbrace{2.00\ \cancel{\text{L}} \times \frac{1000\ \cancel{\text{mL}}}{1\ \cancel{\text{L}}} \times \frac{1.00\ \cancel{\text{g}}}{\cancel{\text{mL}}}}_{\text{Water mass}} \times \underbrace{\frac{1.000\ \text{cal}}{\cancel{\text{g}} \cdot {}^\circ\cancel{\text{C}}}}_{\substack{\text{Specific}\\ \text{heat of water}}} \times \underbrace{18.0\ {}^\circ\text{C}}_{\substack{\text{Temperature}\\ \text{increase}}} = 3.60 \times 10^4\ \text{cal}$$

$$3.60 \times 10^4\ \cancel{\text{cal}} \times \frac{1\ \text{Cal}}{1000\ \cancel{\text{cal}}} = 36.0\ \text{Cal}$$

2.124 To keep all the generated heat inside the unit so that it can warm the water and thereby be measured.

2.125 First, determine the amount of heat released when 2.50 g of wood is burned:

$$0.200\ \cancel{\text{kg}}\ \text{water} \times \frac{1000\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{4.184\ \text{J}}{\cancel{\text{g}} \cdot \cancel{^\circ\text{C}}} \times 6.6\ \cancel{^\circ\text{C}} = 5.5 \times 10^3\ \text{J}$$

$$\frac{5.5 \times 10^3\ \text{J}}{2.50\ \text{g}} = 2.2 \times 10^3\ \text{J per gram of wood}$$

2.126 $$2.00\ \cancel{\text{lb}} \times \frac{453.6\ \cancel{\text{g}}}{1\ \cancel{\text{lb}}} \times \frac{0.449\ \text{J}}{\cancel{\text{g}} \cdot \cancel{^\circ\text{C}}} \times 60.0\ \cancel{^\circ\text{C}} = 2.44 \times 10^4\ \text{J}$$

2.127 You need 2.44×10^4 J of heat energy, and each gram of wood supplies 2.2×10^3 J. Therefore, the mass of wood you need is

$$2.44 \times 10^4\ \cancel{\text{J}} \times \frac{1\ \text{g wood}}{2.2 \times 10^3\ \cancel{\text{J}}} = 11\ \text{g wood}$$

2.128 Manipulate the heat equation to solve for change in temperature and then insert the given data. Because the manipulated equation has a fraction on one side, things get a bit complex, and for this reason it's a good idea to do unit conversions first. Because specific heats are given in the textbook Table 2.5 in J/g · °C convert to grams and joules:

$$2.000\ \cancel{\text{ton}} \times \frac{2000\ \cancel{\text{lb}}}{1\ \cancel{\text{ton}}} \times \frac{453.6\ \text{g}}{1\ \cancel{\text{lb}}} = 1.8144 \times 10^6\ \text{g}$$

(carry the extra significant figure until your final step)

$$8.000 \times 10^6\ \cancel{\text{kJ}} \times \frac{1000\ \text{J}}{1\ \cancel{\text{kJ}}} = 8.000 \times 10^9\ \text{J}$$

These values give a temperature change of

$$\frac{8.000 \times 10^9\ \text{J}}{(0.901\ \text{J/g} \cdot {}^\circ\text{C}) \times 1.8144 \times 10^6\ \text{g}} = 4894\ ^\circ\text{C}$$

Because the initial temperature of the block was 22.0 °C, the block reaches a temperature of 22.0 °C + 4894 °C = 4916 °C. Ouch, hot!

2.129 (c) 1230.0 m has five significant digits, and the converted value must also have five:

$$1230.0\ \cancel{\text{m}} \times \frac{1\ \text{km}}{1000\ \cancel{\text{m}}} = 1.2300\ \text{km}$$

Answer (d) has the correct number of significant digits but the wrong prefix on the unit:

$$1230.0\ \cancel{\text{m}} \times \frac{1000\ \text{mm}}{1\ \cancel{\text{m}}} = 1.2300 \times 10^6\ \text{mm} \neq 1.2300\ \text{mm}$$

2.130 (a) $$(7.98 \times 10^{23}\ \cancel{\mu\text{L}}) \times \frac{1\ \text{L}}{1 \times 10^6\ \cancel{\mu\text{L}}} = 7.98 \times 10^{17}\ \text{L}$$

(b) $$(3.00 \times 10^{-3}\ \cancel{\text{mg}}) \times \frac{1\ \text{g}}{1000\ \cancel{\text{mg}}} = 3.00 \times 10^{-6}\ \text{g}$$

(c) $$(4.21 \times 10^8\ \cancel{\text{mL}}) \times \frac{1\ \cancel{\text{cm}^3}}{1\ \cancel{\text{mL}}} \times \left(\frac{1\ \cancel{\text{m}}}{100\ \cancel{\text{cm}}}\right)^3 \times \frac{264\ \text{gallons}}{1\ \cancel{\text{m}^3}} = 1.11 \times 10^5\ \text{gallons}$$

2.131 $V = \left(\frac{4}{3}\right)\pi r^3 = \left(\frac{4}{3}\right) \times 3.14159 \times (4.00\text{ cm})^3 = 268\text{ cm}^3$

2.132 Because the answer must have grams in it, first convert the given mass to grams:

$$2.5\text{ kg} \times \frac{1000\text{ g}}{1\text{ kg}} = 2.5 \times 10^3\text{ g}$$

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} = \frac{2.5 \times 10^3\text{ g}}{268\text{ cm}^3} = 9.3\text{ g/cm}^3$$

2.133 (b) 0.000 0003 L has only one significant figure, meaning the converted value can have only one. Answer (a) has the correct number of significant figures but the wrong prefix on the unit:

$$3 \times 10^{-6}\text{ L} \times \frac{1000\text{ mL}}{1\text{ L}} = 0.003\text{ mL} \neq 3\text{ mL}$$

2.134 $\frac{60.0\ \cancel{\text{mi}}}{\cancel{\text{h}}} \times \frac{1.61\ \cancel{\text{km}}}{1\ \cancel{\text{mi}}} \times \frac{1000\text{ m}}{1\ \cancel{\text{km}}} \times \frac{1\ \cancel{\text{h}}}{60\ \cancel{\text{min}}} \times \frac{1\ \cancel{\text{min}}}{60\text{ s}} = 26.8\text{ m/s}$

2.135 $\frac{11.0\ \cancel{\text{km}}}{\cancel{\text{L}}} \times \frac{1\text{ mi}}{1.61\ \cancel{\text{km}}} \times \frac{3.79\ \cancel{\text{L}}}{1\text{ gallon}} = 25.9\text{ mi/gallon}$

2.136 (a) $°\text{C} = \left(\frac{5}{9}\right)(72\ °\text{F} - 32) = 22\ °\text{C}$

$\text{K} = 22\ °\text{C} + 273.15 = 295\text{ K}$

(b) $°\text{F} = 32 + \left(\frac{9}{5}\right)(-12\ °\text{C}) = 10\ °\text{F}$

$\text{K} = -12\ °\text{C} + 273.15 = 261\text{ K}$

(c) $°\text{C} = 178\text{ K} - 273.15 = -95\ °\text{C}$

$°\text{F} = 32 + \left(\frac{9}{5}\right)(-95\ °\text{C}) = -139\ °\text{F}$

2.137 1 Calorie = 1000 calories = 1 kilocalorie; 1 calorie = 0.001 Calorie = 1 milliCalorie.

2.138 The volume of the stopper is 37.42 mL − 25.46 mL = 11.96 mL

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} = \frac{16.74\text{ g}}{11.96\text{ mL}} = 1.400\text{ g/mL}$$

2.139 (a) Subtract 32 from both sides and then multiply both sides by 5/9:

$$°\text{F} - 32 = \frac{9}{5}°\text{C} + \cancel{32} - \cancel{32} \Rightarrow °\text{F} - 32 = \frac{9}{5}°\text{C}$$

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

(b) Divide both sides by nR:

$$\frac{PV}{nR} = \frac{nRT}{\cancel{nR}} \Rightarrow \frac{PV}{nR} = T$$

(c) Multiply both sides by λ and divide both sides by E:

$$E \times \lambda = \frac{hc}{\cancel{\lambda}} \times \cancel{\lambda} \Rightarrow E \times \lambda = hc$$

$$\frac{\cancel{E}\lambda}{\cancel{E}} = \frac{hc}{E} \Rightarrow \lambda = \frac{hc}{E}$$

2.140 (a) $2.37 \times 10^2\ \cancel{\text{L}} \times \frac{1000\ \text{mL}}{1\ \cancel{\text{L}}} = 2.7 \times 10^5\ \text{mL}$

(b) $800\ \cancel{\text{kg}} \times \frac{1000\ \text{g}}{1\ \cancel{\text{kg}}} = 8 \times 10^5\ \text{g}$

(c) $0.592\ \cancel{\text{mm}} \times \frac{1\ \text{m}}{1000\ \cancel{\text{mm}}} = 5.92 \times 10^{-4}\ \text{m}$

(d) $8.31\ \cancel{\text{g}} \times \frac{1\ \text{kg}}{1000\ \cancel{\text{g}}} = 8.31 \times 10^{-3}\ \text{kg}$

(e) $9.62 \times 10^{-6}\ \cancel{\text{L}} \times \frac{1 \times 10^6\ \mu\text{L}}{1\ \cancel{\text{L}}} = 9.62\ \mu\text{L}$

(f) $8000\ \cancel{\text{m}} \times \frac{1\ \text{km}}{1000\ \cancel{\text{m}}} = 8\ \text{km}$

(g) $19.3\ \cancel{\text{mg}} \times \frac{1\ \text{g}}{1000\ \cancel{\text{mg}}} = 1.93 \times 10^{-2}\ \text{g}$

(h) $0.003\,45\ \cancel{\text{mL}} \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}} = 3.45 \times 10^{-6}\ \text{L}$

2.141 (a) $\frac{1.34\ \text{g}}{\cancel{\text{L}}} \times \frac{1\ \cancel{\text{L}}}{1000\ \text{mL}} = 1.34 \times 10^{-3}\ \text{g/mL}$

(b) $\frac{1.34\ \cancel{\text{g}}}{\text{L}} \times \frac{1\ \text{kg}}{1000\ \cancel{\text{g}}} = 1.34 \times 10^{-3}\ \text{kg/L}$

(c) $\frac{1.34\ \cancel{\text{g}}}{\cancel{\text{L}}} \times \frac{1\ \text{kg}}{1000\ \cancel{\text{g}}} \times \frac{1\ \cancel{\text{L}}}{1000\ \text{mL}} = 1.34 \times 10^{-6}\ \text{kg/mL}$

2.142 Solve the heat equation for mass by dividing both sides by specific heat and by change in temperature:

Heat = Specific heat × Mass × Change in temperature

$$\text{Mass} = \frac{\text{Heat}}{\text{Specific heat} \times \text{Change in temperature}} = \frac{8.8 \times 10^3\ \cancel{\text{J}}}{2.20\ \cancel{\text{J}}/\text{g} \cdot \cancel{°\text{C}} \times 15\ \cancel{°\text{C}}} = 2.7 \times 10^2\ \text{g}$$

2.143 Multiplying or dividing: the number of significant figures in the answer is determined by which multiplied/divided number has the fewest significant figures. Examples: $725 \times 2.6352 = 1.91 \times 10^3$, $427.45 \div 3.0 = 1.4 \times 10^2$. Adding or subtracting: the number of significant figures in the answer is determined by which added/subtracted number is least certain. Examples: $72\underline{5} + 2.6352 = 72\underline{8}$, $427.45 - 3.\underline{0} = 424.\underline{5}$.

2.144 (a) $23.0\ °\text{C} + 273.15 = 296.2\ \text{K}$

(b) $\left(\frac{5}{9}\right)(98.6\ °\text{F} - 32) = 37.0\ °\text{C}$

(c) Because there is no direct Fahrenheit/Kelvin relationship, you must convert to °C first:

$$296\text{ K} - 273.15 = 23\text{ °C};\ 32 + \left(\frac{9}{5}\right)(23\text{ °C}) = 73\text{ °F}$$

(d) Again convert to °C first, then to Kelvin: $\left(\frac{5}{9}\right)(32\text{ °F} - 32) = 0\text{ °C};\ 0\text{ °C} + 273.15 = 273\text{ K}$

(e) $523\text{ K} - 273.15 = 250\text{ °C}$

(f) $32 + \left(\frac{9}{5}\right)(38\text{ °C}) = 100\text{ °F}$

2.145 (a) Neither accurate nor precise. The large spread between the highest and lowest values means the set is not precise. The average value $(6.38\text{ g} + 9.23\text{ g} + 4.36\text{ g}) \div 3 = 6.66\text{ g}$ tells you the set is not accurate.

(b) Both accurate (the average value is 8.56 g) and precise (very small spread in the three values).

(c) Accurate (the average is 8.54 g) but not precise (large spread).

(d) Precise (very small spread) but not accurate (average 6.26 g).

2.146 The calorie content is a measure of the heat energy contained in the bread, and that contained heat energy is equal to the heat energy absorbed by the water. The change in temperature of the water is $33.0\text{ °C} - 25.0\text{ °C} = 8.0\text{ °C}$. Therefore, after you have converted the water mass to grams, you have

Heat energy absorbed by water = Specific heat of water × Mass of water × Change in temperature

$$\frac{1.00\text{ cal}}{\cancel{\text{g}}\cdot\cancel{\text{°C}}} \times 1000\ \cancel{\text{g}} \times 80\ \cancel{\text{°C}} = 8.0 \times 10^3\text{ cal}$$

The bread's Calorie content is

$$8.0 \times 10^3\ \cancel{\text{cal}} \times \frac{1\text{ Cal}}{1000\ \cancel{\text{cal}}} = 8.0\text{ Cal}$$

2.147 $\text{Density} = \frac{\text{Mass}}{\text{Volume}}$

$\text{Volume} = 3.0\text{ cm} \times 4.0\text{ cm} \times 5.0\text{ cm} = 60\text{ cm}^3$

$$\text{Density} = \frac{470.0\text{ g}}{60.\text{ cm}^3} = 7.8\text{ g/cm}^3$$

2.148 (a) The student is accurate but not precise. The average of her three numbers is 235 g, making her accurate, but the large high–low spread makes her imprecise (but lucky!).

2.149 $\text{Volume} = 28.10\text{ mL} - 25.00\text{ mL} = 3.10\text{ mL}$

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} = \frac{8.34\text{ g}}{3.10\text{ mL}} = 2.69\text{ g/mL}$$

2.150 Heat energy = Specific heat × Mass × Change in temperature
Change in temperature $= 75.0\text{ °C} - 40.0\text{ °C} = 35.0\text{ °C}$

$$\text{Heat energy} = \frac{0.385\text{ J}}{\cancel{\text{g}}\cdot\cancel{\text{°C}}} \times 454\ \cancel{\text{g}} \times 35.0\ \cancel{\text{°C}} = 6.12 \times 10^3\text{ J}$$

$$(6.12 \times 10^3\ \cancel{\text{J}}) \times \frac{1\text{ kJ}}{1000\ \cancel{\text{J}}} = 6.12\text{ kJ}$$

2.151 The conversion equation is $°F = 32 + \left(\frac{9}{5}\right) °C$. Make approximations the quick way by adding 30 instead of 32 and multiplying by 2 instead of $\frac{9}{5}$. Because $\frac{9}{5}$ (1.8) is only a little less than 2 and 32 is only a little more than 30, the errors introduced by making these two approximations tend to cancel, giving a result close to the actual value. As an example, convert 80.5 °C:

Quick way $80.5 \times 2 = 161$

$161 - 16 = 145$

$145 + 30 = 175\ °F$

Equation $32 + \frac{9}{5}(80.5\ °C) = 177\ °F$

2.152 (a) 2.3×10^7 (b) 2.30×10^7 (c) 2.3000×10^7 (d) $2.30\,000 \times 10^7$ (e) $2.3\,000\,000 \times 10^7$

2.153 Although water contains no calories, the body must expend ("burn") energy to raise the temperature of the ice-cold water from approximately 0 °C to the body temperature of 37 °C.

2.154 With markings every 0.01 mL, the estimated (uncertain) digit is in the thousandths position, making the uncertainty ± 0.001 mL.

2.155 Density = Mass/Volume

Density of A = 200.0 g/25.64 mL = 7.800 g/ml

Density of B = 200.0 g/10.36 mL = 19.31 g/ml

Density of C = 200.0 g/17.54 mL = 11.40 g/ml

Therefore A is iron, B is gold, and C is lead.

2.156 Solve the heat equation for change in temperature and then insert the given values, remembering to first convert kilojoules to joules to agree with the specific heat units of J/g · °C:

$$\text{Change in temperature} = \frac{\text{Heat}}{\text{Specific heat} \times \text{Mass}}$$

$$= \frac{10{,}000\ \cancel{J}}{4.184\ \cancel{J}/\cancel{g} \cdot °C \times 250\ \cancel{g}} = 9.6\ °C$$

Final temperature of water = 23.0 °C + 9.6 °C = 32.6 °C

2.157 (a) The uncertain digit, 8, is in the tenths position, making the uncertainty ± 0.1 m.

(b) The uncertain digit, 6, is in the ten-thousandths position, making the uncertainty ± 0.0001 g.

(c) $\pm 0.001 \times 10^3 = \pm 1$ L

(d) ± 1 cm

(e) No uncertainty because 18 here is an exact number.

2.158 (a) 5.93×10^{-1} (b) 4.39×10^5 (c) 7.40×10^{-5}

(d) 2.35×10^{-1} (e) 8.26×10^1 (f) 5.30×10^{-2}

2.159 A Calorie is 1000 calories.

2.160 Precision refers to the closeness to one another of a series of measurements, and therefore the word can never be used to describe a single measurement.

2.161 (a) When adding/subtracting numbers written in scientific notation, first change all numbers to the same power of 10:

$$\begin{array}{r} 5.03 \times 10^2 \\ +\ 0.81 \times 10^2 \\ \hline 5.84 \times 10^2 \end{array} \quad \text{or} \quad \begin{array}{r} 50.3 \times 10^1 \\ +\ 8.1 \times 10^1 \\ \hline 58.4 \times 10^1 \end{array} = 5.84 \times 10^2$$

(b) 4.4×10^{-1}; only two significant digits because of the 0.53.

(c) 2.01×10^{23}; three significant digits because the 3 is exact, meaning the number of significant digits is determined by the 6.02.

(d) As in part (a), change all numbers to the same power of 10:

$$\begin{array}{r} 3.960 \times 10^3 \\ -0.462 \times 10^3 \\ \hline 3.498 \times 10^3 \end{array} \quad \text{or} \quad \begin{array}{r} 39.60 \times 10^2 \\ -\ 4.62 \times 10^2 \\ \hline 34.98 \times 10^2 \end{array} = 3.498 \times 10^3$$

2.162 The volume is length $\times$ width $\times$ height: $6.0\text{ cm} \times 6.0\text{ cm} \times 6.0\text{ cm} = 216\text{ cm}^3$.

The answer can have only two significant digits, but keep the 216 from your calculator display as you do the unit conversion and then round off:

$$216\ \cancel{\text{cm}^3} \times \left(\frac{1\text{ m}}{100\ \cancel{\text{cm}}}\right)^3 = 2.16 \times 10^{-4}\text{ m}^3 = 2.2 \times 10^{-4}\text{ m}^3$$

2.163 (a) $20\ \cancel{\text{atoms}} \times \dfrac{20.2\text{ atomic mass units}}{1\ \cancel{\text{atom}}} = 404\text{ atomic mass units}$

(b) $20\ \cancel{\text{atoms}} \times \dfrac{20.2\ \cancel{\text{atomic mass units}}}{1\ \cancel{\text{atom}}} \times \dfrac{1.66 \times 10^{-24}\text{ g}}{1\ \cancel{\text{atomic mass units}}} = 6.71 \times 10^{-22}\text{ g}$

(c) $6.022 \times 10^{23}\ \cancel{\text{atoms}} \times \dfrac{20.2\ \cancel{\text{atomic mass units}}}{1\ \cancel{\text{atom}}} \times \dfrac{1.66 \times 10^{-24}\text{ g}}{1\ \cancel{\text{atomic mass units}}} = 20.2\text{ g}$

or

$6.022 \times 10^{23}\ \cancel{\text{atoms}} \times \dfrac{6.71 \times 10^{-22}\text{ g}}{20\ \cancel{\text{atoms}}} = 20.2\text{ g}$

2.164 Ethanol because its specific heat is smaller than that of water. The smaller the specific heat, the greater the temperature rise for a given amount of heat energy added.

2.165 Solve the density equation for volume and then insert the data:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} \qquad \text{Volume} = \frac{\text{Mass}}{\text{Density}}$$

(a) $\dfrac{15.0\text{ g}}{0.997\text{ g/mL}} = 15.0\text{ mL}$

(b) $\dfrac{15.0\text{ g}}{0.917\text{ g/mL}} = 16.4\text{ mL}$

(c) $\dfrac{15.0\text{ g}}{0.7\text{ g/mL}} = 21.4\text{ mL} = 2 \times 10^1\text{ mL}$ (only one significant figure allowed)

(d) $\dfrac{15.0\text{ g}}{11.4\text{ g/mL}} = 1.32\text{ mL}$

(e) $\dfrac{15.0\ \text{g}}{13.6\ \text{g/mL}} = 1.10\ \text{mL}$

(f) $\dfrac{15.0\ \text{g}}{0.00018\ \text{g/mL}} = 8.3 \times 10^4\ \text{mL}$

2.166 (a) Significant. The uncertainty ± 1 tells you the uncertain digit is in the ones position, and this uncertain digit is significant.

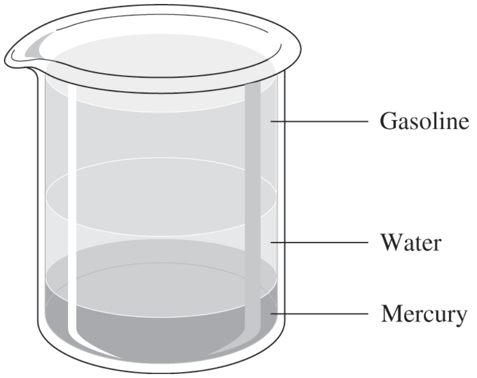
(b) Not significant. The uncertainty ± 10 tells you the uncertain digit is in the tens position.

(c) and (d) Significant because it comes after the decimal point.

(e) It is ambiguous without additional information. This number could be standard notation for either 5.4×10^2 (trailing zero not significant) or 5.40×10^2 (trailing zero significant).

2.167 (a) Six. (b) Three. (c) Two. (d) Four. (e) Four.

2.168 Because it is densest, the mercury is at the bottom. Because it is least dense, the gasoline is at the top.



2.169 (a) True.

(b) False. When adding or subtracting a series of measured values, the number of significant figures in the answer is limited by the least certain measured value. Example: $3724 + 3.2 = 3727$. The 3.2 has the fewest significant figures but is certain to the tenths position. The 3724 is certain only to the ones position, and therefore is the measured value that determines the number of significant figures in the answer—the answer must be uncertain in the ones position. In this example, that restriction means four significant figures are allowed in the answer despite the 3.2.

2.170 (a) $\dfrac{1.04\ \cancel{\text{J}}}{\text{g} \cdot {}^\circ\text{C}} \times \dfrac{1\ \text{cal}}{4.184\ \cancel{\text{J}}} = 0.249\ \text{cal/g} \cdot {}^\circ\text{C}$

(b) $\dfrac{0.84\ \cancel{\text{J}}}{\text{g} \cdot {}^\circ\text{C}} \times \dfrac{1\ \text{cal}}{4.184\ \cancel{\text{J}}} = 0.20\ \text{cal/g} \cdot {}^\circ\text{C}$

2.171 $15\ \cancel{\text{weeks}} \times \dfrac{3\ \cancel{\text{h}}}{\cancel{\text{week}}} \times \dfrac{60\ \cancel{\text{min}}}{1\ \cancel{\text{h}}} \times \dfrac{60\ \cancel{\text{s}}}{1\ \cancel{\text{min}}} \times \dfrac{1000\ \text{ms}}{1\ \cancel{\text{s}}} = 1.6 \times 10^8\ \text{ms}$

2.172 Each tool or device with which a measurement is made has its precision. For instance, a typical ruler has markings every 1 mm (0.1 cm) and one can typically estimate one more position, that is, a multiple of 0.01 cm. So, in any measured value, the last digit written is that uncertain, estimated digit. The uncertainty of the measured value is then indicated by placing a ± 1 in the position of the uncertain digit. For instance, if the measured length is 2.54 cm, one would assume uncertainty in the hundredths position and report it as 2.54 ± 0.01 cm.

2.173 (a) 5.02×10^5 (b) 3.8402×10^{-5} (c) 4.36×10^8
(d) 8.47×10^3 (e) 5.91×10^{-3} (f) 6.58×10^{-1}

2.174 $800\ \text{ft}^2 \times \left(\dfrac{12\ \cancel{\text{inches}}}{1\ \cancel{\text{ft}}}\right)^2 \times \left(\dfrac{1\ \text{m}}{39.37\ \cancel{\text{inches}}}\right)^2 = 74.3\ \text{m}^2$

2.175 $350\ \cancel{\text{Martian dollars}} \times \dfrac{1\ \text{U.S. dollar}}{1.54\ \cancel{\text{Martian dollars}}} = 227\ \text{U.S. dollars}$

2.176 The two blocks contain the same amount of matter. The gram is a unit of mass, and therefore the two blocks have the same mass. *Matter* was defined in Chapter 1 of the textbook as anything that has mass (and occupies space). Thus the two blocks contain the same amount of matter because they have the same mass. (Because lead and gold have different densities, the blocks occupy different *volumes*, but that point is not asked about in the problem.)

2.177 (a) Two. (b) Two. (c) None. (d) Cannot tell. (e) Three.

2.178 (a) 0.189 (b) 793.2 (c) 10^{-14} (d) 0.346

2.179 No. In 580, the decimal point indicates that the trailing zero is significant. In 580, the trailing zero may or may not be significant; there is no way to tell with the value written this way.

2.180 These two temperature units are equal in size because on both the Kelvin scale and the Celsius scale there are 100 units between the freezing point and boiling point of water.

2.181 It is more likely that the student's laboratory technique is bad because his measurements vary widely.

2.182 (a) 0.0179 (b) 0.000 000 008 76 (c) 48,800,000,000 (d) 75.2 (e) 8.37 (f) 4184

2.183 (a) Density = Mass/Volume = $195\ \text{g}/25.0\ \text{cm}^3 = 7.80\ \text{g/cm}^3$

(b) Volume = Mass/Density = $500.0\ \text{g}/7.80\ \text{g/cm}^3 = 64.1\ \text{cm}^3$

(c) The substance floats in mercury because it is less dense than mercury.

2.184 (a) $536\ \cancel{\text{mg}} \times \dfrac{1\ \text{g}}{1000\ \cancel{\text{mg}}} = 0.536\ \text{g} = 5.36 \times 10^{-1}\ \text{g}$

(b) $8.26\ \cancel{\text{dg}} \times \dfrac{1\ \text{g}}{10\ \cancel{\text{dg}}} = 0.826\ \text{g} = 8.26 \times 10^{-1}\ \text{g}$

(c) $0.0057\ \cancel{\mu\text{g}} \times \dfrac{1\ \text{g}}{1 \times 10^6\ \cancel{\mu\text{g}}} = 0.000\,000\,005\,7\ \text{g} = 5.7 \times 10^{-9}\ \text{g}$

(d) $139\ \cancel{\text{kg}} \times \dfrac{1\ \text{g}}{1 \times 10^{-3}\ \cancel{\text{kg}}} = 139{,}000\ \text{g} = 1.39 \times 10^5\ \text{g}$

(e) $836\ \cancel{\text{ng}} \times \dfrac{1\ \text{g}}{1 \times 10^9\ \cancel{\text{ng}}} = 0.000\,000\,836\ \text{g} = 8.36 \times 10^{-7}\ \text{g}$

(f) $0.073\ \cancel{\text{Mg}} \times \dfrac{1\ \text{g}}{1 \times 10^{-6}\ \cancel{\text{Mg}}} = 73{,}000\ \text{g} = 7.3 \times 10^4\ \text{g}$

2.185 (a) $\dfrac{1.3 \times 10^{-3}\ \cancel{\text{g}}}{\cancel{\text{mL}}} \times \dfrac{1\ \text{kg}}{1000\ \text{g}} \times \dfrac{1000\ \cancel{\text{mL}}}{1\ \text{L}} = 1.3 \times 10^{-3}\ \text{kg/L}$

(b) $\dfrac{1.3 \times 10^{-3}\ \cancel{\text{g}}}{\cancel{\text{mL}}} \times \dfrac{1\ \cancel{\text{kg}}}{1000\ \text{g}} \times \dfrac{2.204\ \text{1b}}{1\ \cancel{\text{kg}}} \times \dfrac{1000\ \cancel{\text{mL}}}{1\cancel{\text{L}}} \times \dfrac{1\ \cancel{\text{L}}}{1.057\ \cancel{\text{qt}}} \times \dfrac{4\ \cancel{\text{qt}}}{1\ \text{gallon}} = 1.1 \times 10^{-2}\ \text{1b/gallon}$

2.186 The lemonade in the glass containing the aluminum is cooler because more heat energy has flowed out of the liquid and into the block to change the block's temperature. The block masses are the same, and the block temperature changes are the same. Therefore, it's a matter of looking at the joules that leave each liquid to heat up the blocks. The specific heat of aluminum, $0.901 \text{ J/g} \cdot {}^\circ\text{C}$, is approximately twice that of iron, $0.449 \text{ J/g} \cdot {}^\circ\text{C}$. Thus the amount of heat energy that had to flow from the liquid to heat the aluminum block by 20 °C is approximately twice the amount that had to flow from the liquid to heat the iron block by 20 °C. Having lost twice as much heat energy, the liquid in the glass containing the aluminum is cooler than the liquid in the glass containing the iron.

2.187 (a) 1.0726×10^4. Changing both numbers in the sum from scientific notation to standard notation shows that both are uncertain in the ones position. Therefore, their sum has its uncertainty in the ones position: $9865 + 861 = 10726$.

(b) 4.4×10^{-19}. Only two significant figures allowed because of the 4.5.

(c) 1.471×10^{-18}. The 3.821 restricts the answer to four significant figures.

(d) 9.0618×10^2. Because the 5 is exact, the number of significant figures in the answer is determined by the two numbers of the subtraction. Standard notation shows that both numbers are uncertain in the tenths position, and therefore their difference has five significant figures: $4560.0 - 29.1 = 4530.9$. Dividing this value by the exact number 5 gives an answer having five significant figures.

2.188 $\text{Density} = \dfrac{\text{Mass}}{\text{Volume}}$

$\text{Mass} = \text{Density} \times \text{Volume}$

(a) $\dfrac{11.4 \text{ g}}{\cancel{\text{mL}}} \times 50.0 \cancel{\text{mL}} = 570. \text{ g}$

(b) $\dfrac{0.785 \text{ g}}{\cancel{\text{mL}}} \times 50.0 \cancel{\text{mL}} = 39.3 \text{ g}$

(c) $\dfrac{1.4 \times 10^{-3} \text{ g}}{\cancel{\text{mL}}} \times 50.0 \cancel{\text{mL}} = 0.070 \text{ g}$

(d) $\dfrac{8.4 \times 10^{-5} \text{ g}}{\cancel{\text{mL}}} \times 50.0 \cancel{\text{mL}} = 0.0042 \text{ g}$

(e) $\dfrac{13.6 \text{ g}}{\cancel{\text{mL}}} \times 50.0 \cancel{\text{mL}} = 680. \text{ g}$

(f) $\dfrac{19.3 \text{ g}}{\cancel{\text{mL}}} \times 50.0 \cancel{\text{mL}} = 965 \text{ g}$

2.189 (a) $\dfrac{1 \text{ kg}}{1000 \text{ g}} \qquad \dfrac{1000 \text{ g}}{1 \text{ kg}}$

(b) $\dfrac{1 \text{ g}}{0.001 \text{ kg}} \qquad \dfrac{0.001 \text{ kg}}{1 \text{ g}}$

(c) $\dfrac{1 \text{ yd}}{3 \text{ ft}} \qquad \dfrac{3 \text{ ft}}{1 \text{ yd}}$

(d) $\dfrac{1 \text{ m}}{100 \text{ cm}} \qquad \dfrac{100 \text{ cm}}{1 \text{ m}}$

(e) $\dfrac{1 \text{ cm}}{0.01 \text{ m}} \qquad \dfrac{0.01 \text{ m}}{1 \text{ cm}}$

2.190 $125\ \cancel{\text{lb/in}^2} \times \dfrac{1\ \text{atm}}{14.70\ \cancel{\text{lb/in}^2}} = 8.50\ \text{atm}$

2.191 (a) 4×10^3 (b) 0.37

(c) 10.12. The product of 6.23 and 0.042 is 0.26, only two significant digits but certain to the hundredths position. Add this to 9.86 and you get a sum certain to the hundredths position—legitimately gaining a significant digit.

2.192 Incorrect because 1 ft equals exactly 12 in. by definition. There is no uncertainty.

2.193 (a) Two, determined by the 0.0080.

(b) Three. The $22.1 \times 10^2 = 2210$ is uncertain in the tenths position, meaning the sum must also be: $530 + 2210 = 2740 = 274 \times 10^3$.

(c) Five. The $5.830 \times 10^2 = 583.0$ is uncertain in the tenths position, and the same is true for the $22.100 \times 10^2 = 2210.0$. Therefore, the sum is also uncertain in the tenths position: $583.0 + 2210.0 = 2793.0 = 2.7930 \times 10^3$.

(d) Four, determined by 100.0 and 0.1500.

(e) Two, determined by 0.15.

2.194 $350\ \cancel{\text{Freedonia pounds}} \times \dfrac{1\ \text{U.S. dollar}}{0.690\ \cancel{\text{Freedonia pound}}} = 507\ \text{U.S. dollars}$

2.195 More likely to be precise. Her good laboratory technique will yield measurements that are close to one another (high precision), but the volumes she reads will all be 5 mL higher than the true volume (low accuracy) because of the incorrect markings.

2.196 Because ice is less dense than liquid water, the ice that forms in a lake floats on the liquid water, forming a layer of ice above the liquid water. If ice were denser than liquid water, any ice that formed in a lake would sink to the bottom, and the lake would freeze solid, from the bottom up. It is the fact that ice floats on top that allows a liquid-water environment for fish below, making ice-fishing possible.

2.197 (a) Four. (b) Two. (c) Three. (d) Eight. (e) Four.

2.198 (a) $\dfrac{70\ \cancel{\text{mi}}}{\text{h}} \times \dfrac{1.61\ \text{km}}{1\ \cancel{\text{mi}}} = 1.1 \times 10^2\ \text{km/h}$

(b) $\dfrac{70\ \cancel{\text{mi}}}{1\ \cancel{\text{h}}} \times \dfrac{1.61\ \text{km}}{1\ \cancel{\text{mi}}} \times \dfrac{1\ \cancel{\text{h}}}{60\ \cancel{\text{min}}} \times \dfrac{1\ \cancel{\text{min}}}{60\ \text{s}} = 0.031\ \text{km/s}$

or

$$\dfrac{1.1 \times 10^2\ \text{km}}{\cancel{\text{h}}} \times \dfrac{1\ \cancel{\text{h}}}{3600\ \text{s}} = 0.031\ \text{km/s}$$

(c) $\dfrac{70\ \cancel{\text{mi}}}{1\ \text{h}} \times \dfrac{1.61\ \cancel{\text{km}}}{1\ \cancel{\text{mi}}} \times \dfrac{1000\ \text{m}}{1\ \cancel{\text{km}}} = 1.1 \times 10^5\ \text{m/h}$

or

$$\dfrac{1.1 \times 10^2\ \cancel{\text{km}}}{\text{h}} \times \dfrac{1000\ \text{m}}{1\ \cancel{\text{km}}} = 1.1 \times 10^5\ \text{m/h}$$

(d) $\frac{70\ \cancel{\text{mi}}}{1\ \cancel{\text{h}}} \times \frac{1.61\ \cancel{\text{km}}}{1\ \cancel{\text{mi}}} \times \frac{1000\ \text{m}}{1\ \cancel{\text{km}}} \times \frac{1\ \cancel{\text{h}}}{60\ \cancel{\text{min}}} \times \frac{1\ \cancel{\text{min}}}{60\ \text{s}} = 3.1 \times 10^1\ \text{m/s}$

or

$\frac{1.1 \times 10^5\ \text{m}}{\cancel{\text{h}}} \times \frac{1\ \cancel{\text{h}}}{3600\ \text{s}} = 3.1 \times 10^1\ \text{m/s}$

2.199 To compress the gas means to squeeze a given mass of it into a smaller volume. Compressing therefore increases the density of the gas because a given mass is forced to occupy a smaller volume. The relationship density = mass/volume tells you that, when mass stays constant, density must go up when volume goes down.

2.200 Heat = Specific heat × Mass × Change in temperature

The change in temperature is 37.0 °C − 25.0 °C = 12.0 °C.

(a) $\frac{0.449\ \text{J}}{\cancel{\text{g}} \cdot \cancel{°\text{C}}} \times 50.0\ \cancel{\text{g}} \times 12.0\ \cancel{°\text{C}} = 269\ \text{J}$

(b) $\frac{0.901\ \text{J}}{\cancel{\text{g}} \cdot \cancel{°\text{C}}} \times 50.0\ \cancel{\text{g}} \times 12.0\ \cancel{°\text{C}} = 541\ \text{J}$

(c) $\frac{0.14\ \text{J}}{\cancel{\text{g}} \cdot \cancel{°\text{C}}} \times 50.0\ \cancel{\text{g}} \times 12.0\ \cancel{°\text{C}} = 84\ \text{J}$

(d) $\frac{4.18\ \text{J}}{\cancel{\text{g}} \cdot \cancel{°\text{C}}} \times 50.0\ \cancel{\text{g}} \times 12.0\ \cancel{°\text{C}} = 2.51 \times 10^3\ \text{J}$

2.201 Water undergoes the most gradual temperature change because it has the highest specific heat. Mercury undergoes the fastest temperature change because it has the lowest specific heat.

2.202 Shooter A is more precise (there is little spread of data—all the arrows hit the same mark). Shooter B is more accurate (all three arrows come closer to a bulls eye than Shooter A).

2.203 Bar A represents a joule and bar B a calorie. A joule is about one-fourth of a calorie (4.184 J = 1 cal).

2.204 3.5000×10^9

2.205 40.2×10^{-6} (or 4.02×10^{-5}) m; 40.2×10^3 (4.02×10^4) nm

2.206 The mass/centimeter of an ordinary Twinkie is 1.0 g/cm. After converting feet to centimeters we find that a 6.4 ft Twinkie is 2.0×10^1 cm long.

The mass of the 6.4 ft. Twinkie is therefore 2.0×10^1 g. After converting 2.0×10^1 g to tons, we find that the Twinkie weighs 2.2×10^{-4} tons.

2.207 Heat = Specific heat of water × Mass of water in calorimeter × Change in water temperature

$\text{Heat} = \frac{4.184\ \text{J}}{\cancel{\text{g}} \cdot °\text{C}} \times 1.00 \times 10^4\ \cancel{\text{g}} \times 20.0\ \cancel{°\text{C}} = 8.37 \times 10^5\ \text{J or } 8.37 \times 10^2\ \text{kJ of energy.}$

$\text{Therefore the kJ/g} = \frac{8.37 \times 10^2\ \text{kJ}}{10.6\ \text{g}} = 79.0\ \text{kJ/g}$

2.208 No, he is not right. There is some uncertainty in any measurement. The uncertainty in the mass obtained by using a digital scale lies in the last digit, making the mass 2.635 ± 0.001 g.

2.209 The mercury would have the smallest volume. It is by far the most dense, meaning that 50 g would fit into a smaller volume than any of the other substances.

2.210 $325 \times 10^{-3}\text{ g} \times \dfrac{1\text{ mL}}{5000 \times 10^{-6}\text{ g}} = 65\text{ mL}$

Answers to Concept Questions:

1. (a), **2.** (c), **3.** (c), **4.** (c), **5.** (c), **6.** (a), **7.** (c), **8.** (c), **9.** (b), **10.** (b)