SOLUTIONS TO EXERCISES

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TWELFTH EDITION

CHEMISTRY

THE CENTRAL SCIENCE

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Introduction

Chemistry: The Central Science, 12th edition, contains more than 2600 end-of-chapter exercises. Considerable attention has been given to these exercises because one of the best ways for students to master chemistry is by solving problems. Grouping the exercises according to subject matter is intended to aid the student in selecting and recognizing particular types of problems. Within each subject matter group, similar problems are arranged in pairs. This provides the student with an opportunity to reinforce a particular kind of problem. There are also a substantial number of general exercises in each chapter to supplement those grouped by topic. Integrative exercises, which require students to integrate concepts from several chapters, are a continuing feature of the 12th edition. Answers to the odd numbered topical exercises plus selected general and integrative exercises, about 1200 in all, are provided in the text. These appendix answers help to make the text a useful self-contained vehicle for learning.

This manual, **Solutions to Exercises in Chemistry: The Central Science, 12th edition,** was written to enhance the end-of-chapter exercises by providing documented solutions. The manual assists the instructor by saving time spent generating solutions for assigned problem sets and aids the student by offering a convenient independent source to check their understanding of the material. Most solutions have been worked in the same detail as the in-chapter sample exercises to help guide students in their studies.

To reinforce the *'Analyze, Plan, Solve, Check'* problem-solving method used extensively in the text, this strategy has also been incorporated into the Solution Manual. Solutions to most red paired exercises and selected Additional and Integrative exercises feature this four-step approach. We strongly encourage students to master this powerful and totally general method.

When using this manual, keep in mind that the numerical result of any calculation is influenced by the precision of the numbers used in the calculation. In this manual, for example, atomic masses and physical constants are typically expressed to four significant figures, or at least as precisely as the data given in the problem. If students use slightly different values to solve problems, their answers will differ slightly from those listed in the appendix of the text or this manual. This is a normal and a common occurrence when comparing results from different calculations or experiments.

Rounding methods are another source of differences between calculated values. In this manual, when a solution is given in steps, intermediate results will be rounded to the correct number of significant figures; however, unrounded numbers will be used in subsequent calculations. By following this scheme, calculators need not be cleared to re-enter rounded intermediate results in the middle of a calculation sequence. The final answer will appear with the correct number of significant figures. This may result in a small discrepancy in the last significant digit between student-calculated answers and those given in this manual. Variations due to rounding can occur in any analysis of numerical data.

The first step in checking your solution and resolving differences between your answer and the listed value is to look for similarities and differences in problem-solving methods. Ultimately, resolving the small numerical differences described above is less important than understanding the general method for solving a problem. The goal of this manual is to provide a reference for sound and consistent problem-solving methods in addition to accurate answers to text exercises.

Extraordinary efforts have been made to keep this manual as error-free as possible. All exercises were worked and proof-read by at least three chemists to ensure clarity in methods and accuracy in mathematics. The work and advice of Ms. Rene Musto, Ms. Kate Vigor, Dr. Christopher Musto and Dr. Timothy Kucharski have been invaluable to this project. In any written work as technically challenging as this manual, typos and errors inevitably creep in, despite our combined efforts. Please help us find and eliminate them. We hope that both instructors and students will find this manual accurate, helpful and instructive.

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1 Introduction: Matter and Measurement

Visualizing Concepts

1.1 *Pure elements* contain only one kind of atom. Atoms can be present singly or as tightly bound groups called molecules. *Compounds* contain two or more kinds of atoms bound tightly into molecules. *Mixtures* contain more than one kind of atom and/or molecule, not bound into discrete particles.

(a) pure element: i

(b) mixture of elements: v, vi

(c) pure compound: iv

(d) mixture of an element and a compound: ii, iii

1.2 After a *physical change,* the identities of the substances involved are the same as their identities before the change. That is, molecules retain their original composition. During a *chemical change,* at least one new substance is produced; rearrangement of atoms into new molecules occurs.

The diagram represents a chemical change, because the molecules after the change are different than the molecules before the change.

1.3 To brew a cup of coffee, begin with ground coffee beans, a heterogeneous mixture, and water, a pure substance. Hot water contacts the coffee grounds and dissolves components of the coffee bean that are water-soluble. This creates a new heterogeneous mixture of undissolved coffee bean solids and liquid coffee solution; this mixture is separated by filtration. Undissolved grounds are left on the filter paper and liquid coffee, itself a homogeneous mixture, drips into the container below.

Overall, two separations occur. Chemical differences among the components of the coffee bean allow certain compounds to dissolve in water, while other components remain insoluble. This kind of separation based on solubility differences is called *extraction*. The insoluble grounds are then separated from the coffee solution by *filtration*.

1.4 (a) time (b) mass (c) temperature (d) area (e) length

(f) area (g) temperature (h) density (i) volume

1.5 Density is the ratio of mass to volume. For a sphere, size is like volume; both are determined by the radius of the sphere.

(a) For spheres of the same size or volume, the denominator of the density relationship is the same. The denser the sphere, the heavier it is. A list from lightest to heaviest is in order of increasing density and mass. The aluminum sphere (density = 2.70 g/cm^3) is lightest, then nickel (density = 8.90 g/cm^3), then silver (density = 10.409 g/cm^3).

(b) For spheres of equal mass, the numerator of the density relationship is the same. The denser the sphere, the smaller its volume or size. A list from smallest to largest is in order of decreasing density. The platinum sphere (density = 21.45 g/cm^3) is smallest, then gold (density = 19.30 g/cm^3), then lead (density = 11.35 g/cm^3).

1.6 Measurements (darts) that are close to each other are *precise*. Measurements that are close to the "true value" (the bull's eye) are *accurate*.

(a) Figure ii represents data that are both accurate and precise. The darts are close to the bull's eye and each other.

(b) Figure i represents data that are precise but inaccurate. The darts are near each other but their center point (average value) is far from the bull's eye.

(c) Figure iii represents data that are imprecise but their average value is accurate. The darts are far from each other, but their average value, or geometric center point, is close to the bull's eye.

1.7 (a) 7.5 cm. There are two significant figures in this measurement; the number of cm can be read precisely, but there is some estimating (uncertainty) required to read tenths of a centimeter. Listing two significant figures is consistent with the convention that measured quantities are reported so that there is uncertainty in only the last digit.

(b) The speed is 72 mi/hr (inner scale, two significant figures) or 115 km/hr (outer scale, three significant figures). Both scales are read with certainty in the "hundreds" and "tens" place, and some uncertainty in the "ones" place. The km/hr speed has one more significant figure because its magnitude is in the hundreds.

1.8 (a) Volume = length × width × height. Because the operation is multiplication, the dimension with fewest significant figures (sig figs) determines the number of sig figs in the result. The dimension "2.5 cm" has 2 sig figs, so the volume is reported with 2 sig figs.

(b) Density = mass/volume. Because the operation is division, again the datum with fewer significant figures determines the number of sig figs in the result. While mass, 104.72 g, has 5 sig figs, volume [from (a)] has 2 sig figs, so density is also reported to 2 sig figs.

1.9 When converting units, arrange the conversion factor so that the given unit cancels and the desired unit is in the correct position. For example, suppose a quantity is expressed in terms of centimeters, but the desired result is expressed in inches. If the given unit has 'cm' in the numerator, then the conversion factor must have 'cm' in its denominator. However, if the original unit has 'cm' in the denominator, the conversion factor must have 'cm' in the numerator. Ideally, this will lead to the desired units in the appropriate location, numerator or denominator. However, the inverse of the answer can be taken when necessary.

1.10 Given: m/s Find: mi/hr. Both the given and desired units have distance in the numerator and time in the denominator. Use appropriate conversion factors to change 'm' to 'mi' in the numerator and 's' to 'hr' in the denominator.

m/s → (use 1 km/10^3 m) → km/s → (use 1 mi/1.6093 km) → mi/s → (use 60 sec/1 min) → mi/min → (use 60 min/1 hr) → mi/hr

Classification and Properties of Matter (sections 1.2 and 1.3)

1.11 (a) heterogeneous mixture

(b) homogeneous mixture (If there are undissolved particles, such as sand or decaying plants, the mixture is heterogeneous.)

(c) pure substance

(d) pure substance

1.12 (a) homogeneous mixture

(b) heterogeneous mixture (particles in liquid)

(c) pure substance

(d) heterogeneous mixture

1.13 (a) S (b) Au (c) K (d) Cl (e) Cu (f) uranium

(g) nickel (h) sodium (i) aluminum (j) silicon

1.14 (a) C (b) N (c) Ti (d) Zn (e) Fe (f) phosphorus

(g) calcium (h) helium (i) lead (j) silver

1.15 $A(s) \rightarrow B(s) + C(g)$

When solid carbon is burned in excess oxygen gas, the two elements combine to form a gaseous compound, carbon dioxide. Clearly substance C is this compound. Since C is produced when A is heated in the absence of oxygen (from air), both the carbon and oxygen in C must have been present in A originally. A is, therefore, a compound composed of two or more elements chemically combined. Without more information on the chemical or physical properties of B, we cannot determine absolutely whether it is an element or a compound. However, few if any elements exist as white solids, so B is probably also a compound.

1.16 Gold, Au, is an element and "fool's gold", FeS_2, is a compound; both are solids and pure substances. Take advantage of differences in physical and or chemical properties between the two substances. Density and melting point measurements are often used to identify solids. For these two substances, melting points are very high, but densities are easy to measure. Gold is much denser than "fool's gold". Gold is much less chemically reactive than FeS_2, so relative reactivity with acids and bases can be observed. Of these experiments, density measurement is the most definitive and does not destroy the sample. (Note that neither substance is attracted to a magnet, so this test will not identify the gold.)

1.17 *Physical properties*: silvery white (color); lustrous; melting point = 649°C; boiling point = 1105°C; density at 20°C = 1.738 g/cm^3; pounded into sheets (malleable); drawn into wires (ductile); good conductor. *Chemical properties*: burns in air to give intense white light; reacts with Cl_2 to produce brittle white solid.

1.18 *Physical properties*: silver-gray (color); melting point = 420°C; hardness = 2.5 Mohs; density = 7.13 g/cm^3 at 25°C. *Chemical properties*: metal; reacts with sulfuric acid to produce hydrogen gas; reacts slowly with oxygen at elevated temperatures to produce ZnO.

1.19 (a) chemical (b) physical (c) physical (d) chemical (e) chemical

1.20 (a) chemical

(b) physical

(c) physical (The production of H_2O is a chemical change, but its *condensation* is a physical change.)

(d) physical (The production of soot is a chemical change, but its *deposition* is a physical change.)

1.21 (a) Take advantage of the different water solubilities of the two solids. Add water to dissolve the sugar; filter this mixture, collecting the sand on the filter paper and the sugar water in the flask. Evaporate water from the flask to recover solid sugar.

(b) Take advantage of the different solubilities and densities of the two liquids. Allow the mixture to settle so that there are two distinct layers. Vinegar (a water solution) is denser and on the bottom; oil (the organic layer) is less dense and on top. Carefully pour off most of the top layer. After the layers reform, use a dropper to remove any remaining oil. Vinegar is in the original vessel and oil is in a second container.

1.22 First heat the liquid in each beaker to 100°C to evaporate the water. The beaker with no residue contained pure water. The other two beakers have a solid, white residue. Measure the melting point of each solid. Sugar has a much lower melting point than salt, so the beaker with the lower-melting residue contained sugar water and that with the higher-melting residue contained salt water. (If confirmation is required, measure the densities of the two white residues.)

Units and Measurement (section 1.4)

1.23 (a) 1×10^{-1} (b) 1×10^{-2} (c) 1×10^{-15} (d) 1×10^{-6} (e) 1×10^{6}

(f) 1×10^{3} (g) 1×10^{-9} (h) 1×10^{-3} (i) 1×10^{-12}

1.24 (a) $2.3 \times 10^{-10}\ \text{L} \times \dfrac{1\ \text{nL}}{1 \times 10^{-9}\ \text{L}} = 0.23\ \text{nL}$

(b) $4.7 \times 10^{-6}\,g \times \frac{1\,\mu g}{1 \times 10^{-6}\,g} = 4.7\,\mu g$

(c) $1.85 \times 10^{-12}\,m \times \frac{1\,pm}{1 \times 10^{-12}\,m} = 1.85\,pm$

(d) $16.7 \times 10^{6}\,s \times \frac{1\,Ms}{1 \times 10^{6}\,s} = 16.7\,Ms$

(e) $15.7 \times 10^{3}\,g \times \frac{1\,kg}{1 \times 10^{3}\,g} = 15.7\,kg$

(f) $1.34 \times 10^{-3}\,m \times \frac{1\,mm}{1 \times 10^{-3}\,m} = 1.34\,mm$

(g) $1.84 \times 10^{2}\,cm \times \frac{1\,m}{1 \times 10^{2}\,cm} = 1.84\,m$

1.25 (a) °C = 5/9 (°F – 32°); 5/9 (72 – 32) = 22°C

(b) °F = 9/5 (°C) + 32°; 9/5 (216.7) + 32 = 422.1°F

(c) K = °C + 273.15; 233°C + 273.15 = 506 K

(d) °C = 315 K – 273.15 = 41.85 = 42°C; °F = 9/5 (41.85°C) + 32 = 107°F

(e) °C = 5/9 (°F – 32°); 5/9 (2500 – 32) = 1371°C; K = 1371°C + 273.15 = 1644 K (assuming 2500 °F has 4 sig figs)

(f) °C = 0 K – 273.15 = –273.15°C; °F = 9/5 (–273.15°C) + 32 = –459.67°F (assuming 0 K has infinite sig figs)

1.26 (a) °C = 5/9 (87°F – 32°) = 31°C

(b) K = 25°C + 273.15 = 298 K; °F = 9/5 (25°C) + 32 = 77°F

(c) °C = 5/9 (400°F – 32°) = 204.444 = 204°C

K = °C + 273.15 = 204.444°C + 273.15 = 478 K

(d) °C = 77 K – 273.15 = –196.15 = –196°C; °F = 9/5 (–196.15°C) + 32 = –321°F

1.27 (a) $\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{40.55\,g}{25.0\,mL} = 1.62\,g/mL \text{ or } 1.62\,g/cm^3$

(The units cm^3 and mL will be used interchangeably in this manual.)

Tetrachloroethylene, 1.62 g/mL, is more dense than water, 1.00 g/mL; tetrachloroethylene will sink rather than float on water.

(b) $25.0\,cm^3 \times 0.469\,\frac{g}{cm^3} = 11.7\,g$

1.28 (a) volume = length^3 (cm^3); density = mass/volume (g/cm^3)

volume = $(1.500)^3\,cm^3 = 3.375\,cm^3$

$\text{density} = \frac{76.31\,g}{3.375\,cm^3} = 22.61\,g/cm^3$ osmium

(b) $125.0\ \text{mL} \times \frac{1\ \text{cm}^3}{1\ \text{mL}} \times \frac{4.51\ \text{g}}{1\ \text{cm}^3} = 563.75 = 564\ \text{g titanium}$

(c) $0.1500\ \text{L} \times \frac{1\ \text{mL}}{1 \times 10^{-3}\ \text{L}} \times \frac{0.8787\ \text{g}}{1\ \text{mL}} = 131.8\ \text{g benzene}$

1.29 (a) $\text{density} = \frac{38.5\ \text{g}}{45\ \text{mL}} = 0.86\ \text{g/mL}$

The substance is probably toluene, density = 0.866 g/mL.

(b) $45.0\ \text{g} \times \frac{1\ \text{mL}}{1.114\ \text{g}} = 40.4\ \text{mL ethylene glycol}$

(c) $(5.00)^3\ \text{cm}^3 \times \frac{8.90\ \text{g}}{1\ \text{cm}^3} = 1.11 \times 10^3\ \text{g (1.11 kg) nickel}$

1.30 (a) $\frac{21.95\ \text{g}}{25.0\ \text{mL}} = 0.878\ \text{g/mL}$

The tabulated value has four significant figures, while the experimental value has three. The tabulated value rounded to three figures is 0.879. The values agree within one in the last significant figure of the experimental value; the two results agree. The liquid could be benzene.

(b) $15.0\ \text{g} \times \frac{1\ \text{mL}}{0.7781\ \text{g}} = 19.3\ \text{mL cyclohexane}$

(c) $r = d/2 = 5.0\ \text{cm}/2 = 2.5\ \text{cm}$

$V = 4/3\ \pi r^3 = 4/3 \times \pi \times (2.5)^3\ \text{cm}^3 = 65.4498 = 65\ \text{cm}^3$

$65.4498\ \text{cm}^3 \times \frac{11.34\ \text{g}}{\text{cm}^3} = 7.4 \times 10^2\ \text{g}$

(The answer has two significant figures because the diameter had only two significant figures.)

Note: This is the first exercise where "intermediate rounding" occurs. In this manual, when a solution is given in steps, the intermediate result will be rounded to the correct number of significant figures. However, the **unrounded** number will be used in subsequent calculations. The final answer will appear with the correct number of significant figures. That is, calculators need not be cleared and new numbers entered in the middle of a calculation sequence. This may result in a small discrepancy in the last significant digit between student-calculated answers and those given in the manual. These variations occur in any analysis of numerical data.

For example, in this exercise the volume of the sphere, $65.4498\ \text{cm}^3$, is rounded to $65\ \text{cm}^3$, but 65.4498 is retained in the subsequent calculation of mass, 7.4×10^2 g. In this case, $65\ \text{cm}^3 \times 11.34\ \text{g/cm}^3$ also yields 7.4×10^2 g. In other exercises, the correctly rounded results of the two methods may not be identical.

1.31 $31 \text{ billion tons} \times \frac{1 \times 10^9 \text{ tons}}{1 \text{ billion tons}} \times \frac{2000 \text{ lb}}{1 \text{ ton}} \times \frac{453.59 \text{ g}}{1 \text{ lb}} = 2.8 \times 10^{16} \text{ g}$

The metric prefix for 1×10^{15} is peta, abbreviated P.

$$2.8 \times 10^{16} \text{ g} \times \frac{1 \text{ Pg}}{1 \times 10^{15} \text{ g}} = 28 \text{ Pg}$$

1.32 (a) The wafers have the same diameter as the boule, so the question becomes 'how many 0.75 mm wafers can be cut from the 2 m boule?'

$$\frac{2.0 \text{ m}}{\text{boule}} \times \frac{1 \text{ mm}}{1 \times 10^{-3} \text{ m}} \times \frac{1 \text{ wafer}}{0.75 \text{mm}} = 2667 = 2.7 \times 10^3 \text{ wafers}$$

(b) Calculate the volume of the wafer in cm^3. $V = \pi r^2 h$

$$r = \frac{d}{2} = \frac{300 \text{ mm}}{2} \times \frac{1 \text{ cm}}{10 \text{ mm}} = 15 \text{ cm}; \quad h = 0.75 \text{ mm} \times \frac{1 \text{ cm}}{10 \text{ mm}} = 7.5 \times 10^{-2} \text{ cm}$$

$$V = \pi r^2 h = \pi (15 \text{ cm})^2 (7.5 \times 10^{-2} \text{ cm}) = 53.0144 = 53 \text{ cm}^3$$

Density = mass / V; mass = density × V

$$\frac{2.33 \text{ g}}{\text{cm}^3} \times 53.0144 \text{ cm}^3 = 123.52 = 1.2 \times 10^2 \text{ g}$$

Uncertainty in Measurement (section 1.5)

1.33 Exact: (c), (d), and (f) (All others depend on measurements and standards that have margins of error, e.g., the length of a week as defined by the earth's rotation.)

1.34 Exact: (b), (e) (The number of students is exact on any given day.)

1.35 (a) 3 (b) 2 (c) 5 (d) 3 (e) 5 (f) 1 [See Sample Exercise 1.6 (c)]

1.36 (a) 4 (b) 3 (c) 4 (d) 5 (e) 6 (f) 2

1.37 (a) 1.025×10^2 (b) 6.570×10^5 (c) 8.543×10^{-3}

(d) 2.579×10^{-4} (e) -3.572×10^{-2}

1.38 (a) 7.93×10^3 mi (b) 4.001×10^4 km

1.39 (a) 14.3505 + 2.65 = 17.0005 = 17.00 (For addition and subtraction, the minimum number of decimal places, here two, determines decimal places in the result.)

(b) 952.7 − 140.7389 = 812.0

(c) $(3.29 \times 10^4)(0.2501) = 8.23 \times 10^3$ (For multiplication and division, the minimum number of significant figures, here three, determines sig figs in the result.)

(d) $0.0588/0.677 = 8.69 \times 10^{-2}$

1.40 (a) [320.5 - 6104.5/2.3] = -2.3×10^3 (The intermediate result has two significant figures, so only the thousand and hundred places in the answer are significant.)

(b) $[285.3 \times 10^5 - 0.01200 \times 10^5] \times 2.8954 = 8.260 \times 10^7$ (Since subtraction depends on decimal places, both numbers must have the same exponent to determine decimal places/sig figs. The intermediate result has 1 decimal place and 4 sig figs, so the answer has 4 sig figs.)

(c) $(0.0045 \times 20{,}000.0)$ + (2813×12) = 3.4×10^4
2 sig figs /0 dec pl 2 sig figs / first 2 digits

(d) 863 × [1255 - (3.45×108)] = 7.62×10^5
(3 sig figs /0 dec pl)
3 sig figs × [0 dec pl/3 sig figs] = 3 sig figs

1.41 The mass 21.427 g has 5 significant figures.

1.42 The volume in the graduated cylinder is 19.5 mL. Liquid volumes are read at the bottom of the meniscus, so the volume is slightly less than 20 mL. Volumes in this cylinder can be read with certainty to 1 mL, and with some uncertainty to 0.1 mL, so this measurement has 3 sig figs.

Dimensional Analysis (section 1.6)

1.43 In each conversion factor, the old unit appears in the denominator, so it cancels, and the new unit appears in the numerator.

(a) $\text{mm} \rightarrow \text{nm}: \dfrac{1 \times 10^{-3}\ \text{m}}{1\ \text{mm}} \times \dfrac{1\ \text{nm}}{1 \times 10^{-9}\ \text{m}} = 1 \times 10^6\ \text{nm/mm}$

(b) $\text{mg} \rightarrow \text{kg}: \dfrac{1 \times 10^{-3}\ \text{g}}{1\ \text{mg}} \times \dfrac{1\ \text{kg}}{1000\ \text{g}} = 1 \times 10^{-6}\ \text{kg/mg}$

(c) $\text{km} \rightarrow \text{ft}: \dfrac{1000\ \text{m}}{1\ \text{km}} \times \dfrac{1\ \text{cm}}{1 \times 10^{-2}\ \text{m}} \times \dfrac{1\ \text{in}}{2.54\ \text{cm}} \times \dfrac{1\ \text{ft}}{12\ \text{in}} = 3.28 \times 10^3\ \text{km/ft}$

(d) $\text{in}^3 \rightarrow \text{cm}^3: \dfrac{(2.54)^3\ \text{cm}^3}{1^3\ \text{in}^3} = 16.4\ \text{cm}^3/\text{in}^3$

1.44 In each conversion factor, the old unit appears in the denominator, so it cancels, and the new unit appears in the numerator.

(a) $\mu\text{m} \rightarrow \text{mm}: \dfrac{1 \times 10^{-6}\ \text{m}}{1\ \mu\text{m}} \times \dfrac{1\ \text{mm}}{1 \times 10^{-3}\ \text{m}} = 1 \times 10^{-3}\ \text{mm}/\mu\text{m}$

(b) $\text{ms} \rightarrow \text{ns}: \dfrac{1 \times 10^{-3}\ \text{s}}{1\ \text{ms}} \times \dfrac{1\ \text{ns}}{1 \times 10^{-9}\ \text{s}} = 1 \times 10^6\ \text{ns/ms}$

(c) mi → km: 1.6093 km/mi

(d) $\text{ft}^3 \rightarrow \text{L}: \dfrac{(12)^3\ \text{in}^3}{1\ \text{ft}^3} \times \dfrac{(2.54)^3\ \text{cm}^3}{1\ \text{in}^3} \times \dfrac{1\ \text{L}}{1000\ \text{cm}^3} = 28.3\ \text{L/ft}^3$

1.45 (a) $\frac{15.2\text{ m}}{\text{s}} \times \frac{1\text{ km}}{1000\text{ m}} \times \frac{60\text{ s}}{1\text{ min}} \times \frac{60\text{ min}}{1\text{ hr}} = 54.7\text{ km/hr}$

(b) $5.0 \times 10^3\text{ L} \times \frac{1\text{ gal}}{3.7854\text{ L}} = 1.3 \times 10^3\text{ gal}$

(c) $151\text{ ft} \times \frac{1\text{ yd}}{3\text{ ft}} \times \frac{1\text{ m}}{1.0936\text{ yd}} = 46.025 = 46.0\text{ m}$

(d) $\frac{60.0\text{ cm}}{\text{d}} \times \frac{1\text{ in}}{2.54\text{ cm}} \times \frac{1\text{ d}}{24\text{ hr}} = 0.984\text{ in/hr}$

1.46 (a) $\frac{2.998 \times 10^8\text{ m}}{\text{s}} \times \frac{1\text{ km}}{1000\text{ m}} \times \frac{1\text{ mi}}{1.6093\text{ km}} \times \frac{60\text{ s}}{1\text{ min}} \times \frac{60\text{ min}}{1\text{ hr}} = 6.707 \times 10^8\text{ mi/hr}$

(b) $1454\text{ ft} \times \frac{1\text{ yd}}{3\text{ ft}} \times \frac{1\text{ m}}{1.0936\text{ yd}} = 443.18 = 443.2\text{ m}$

(c) $3{,}666{,}500\text{ m}^3 \times \frac{1^3\text{ dm}^3}{(1 \times 10^{-1})^3\text{ m}^3} \times \frac{1\text{ L}}{1\text{ dm}^3} = 3.6665 \times 10^9\text{ L}$

(d) $\frac{242\text{ mg cholesterol}}{100\text{ mL blood}} \times \frac{1\text{ mL}}{1 \times 10^{-3}\text{ L}} \times 5.2\text{ L} \times \frac{1 \times 10^{-3}\text{ g}}{1\text{ mg}} = 12.58 = 13\text{ g cholesterol}$

1.47 (a) $5.00\text{ days} \times \frac{24\text{ hr}}{1\text{ day}} \times \frac{60\text{ min}}{1\text{ hr}} \times \frac{60\text{ s}}{1\text{ min}} = 4.32 \times 10^5\text{ s}$

(b) $0.0550\text{ mi} \times \frac{1.6093\text{ km}}{\text{mi}} \times \frac{1000\text{ m}}{1\text{ km}} = 88.5\text{ m}$

(c) $\frac{\$1.89}{\text{gal}} \times \frac{1\text{ gal}}{3.7854\text{ L}} = \frac{\$0.499}{\text{L}}$

(d) $\frac{0.510\text{ in}}{\text{ms}} \times \frac{2.54\text{ cm}}{1\text{ in}} \times \frac{1 \times 10^{-2}\text{ m}}{1\text{ cm}} \times \frac{1\text{ km}}{1000\text{ m}} \times \frac{1\text{ ms}}{1 \times 10^{-3}\text{ s}} \times \frac{60\text{ s}}{1\text{ min}} \times \frac{60\text{ min}}{1\text{ hr}} = 46.6\ \frac{\text{km}}{\text{hr}}$

Estimate: $0.5 \times 2.5 = 1.25$; $1.25 \times 0.01 \approx 0.01$; $0.01 \times 60 \times 60 \approx 36\text{ km/hr}$

(e) $\frac{22.50\text{ gal}}{\text{min}} \times \frac{3.7854\text{ L}}{\text{gal}} \times \frac{1\text{ min}}{60\text{ s}} = 1.41953 = 1.420\text{ L/s}$

Estimate: $20 \times 4 = 80$; $80/60 \approx 1.3\text{ L/s}$

(f) $0.02500\text{ ft}^3 \times \frac{12^3\text{ in}^3}{1\text{ ft}^3} \times \frac{2.54^3\text{ cm}^3}{1\text{ in}^3} = 707.9\text{ cm}^3$

Estimate: $10^3 = 1000$; $3^3 = 27$; $1000 \times 27 = 27{,}000$; $27{,}000/0.04 \approx 700\text{ cm}^3$

1.48 (a) $0.105\text{ in} \times \frac{2.54\text{ cm}}{\text{in}} \times \frac{1 \times 10^{-2}\text{ m}}{\text{cm}} \times \frac{1\text{ mm}}{1 \times 10^{-3}\text{ m}} = 2.667 = 2.67\text{ mm}$

(b) $0.650\text{ qt} \times \frac{1\text{ L}}{1.057\text{ qt}} \times \frac{1\text{ mL}}{1 \times 10^{-3}\text{ L}} = 614.94 = 615\text{ mL}$

(c) $\frac{8.75\ \mu\text{m}}{\text{s}} \times \frac{1 \times 10^{-6}\text{ m}}{1\ \mu\text{m}} \times \frac{1\text{ km}}{1 \times 10^3\text{ m}} \times \frac{60\text{ s}}{1\text{ min}} \times \frac{60\text{ min}}{1\text{ hr}} = 3.15 \times 10^{-5}\text{ km/hr}$

(d) $1.955\ m^3 \times \frac{(1.0936)^3\ yd^3}{1\ m^3} = 2.55695 = 2.557\ yd^3$

(e) $\frac{\$3.99}{lb} \times \frac{2.205\ lb}{1\ kg} = 8.798 = \$8.80/kg$

(f) $\frac{8.75\ lb}{ft^3} \times \frac{453.59\ g}{1\ lb} \times \frac{1\ ft^3}{12^3\ in^3} \times \frac{1\ in^3}{2.54^3\ cm^3} \times \frac{1\ cm^3}{1\ mL} = 0.140\ g/mL$

1.49 (a) $31\ gal \times \frac{4\ qt}{1\ gal} \times \frac{1\ L}{1.057\ qt} = 1.2 \times 10^2\ L$

Estimate: $(30 \times 4)/1 \approx 120$ L

(b) $\frac{6\ mg}{kg\ (body)} \times \frac{1\ kg}{2.205\ lb} \times 185\ lb = 5 \times 10^2\ mg$

Estimate: $6/2 = 3; 3 \times 180 = 540$ mg

(c) $\frac{400\ km}{47.3\ L} \times \frac{1\ mi}{1.6093\ km} \times \frac{1\ L}{1.057\ qt} \times \frac{4\ qt}{1\ gal} = \frac{19.9\ mi}{gal}$

(2×10^1 mi/gal for 1 sig fig)

Estimate: $400/50 = 8; 8/1.6 = 5; 5/1 = 5; 5 \times 4 \approx 20$ mi/gal

(d) $\frac{50\ cups}{1\ lb} \times \frac{1\ qt}{4\ cups} \times \frac{1\ L}{1.057\ qt} \times \frac{1000\ mL}{1\ L} \times \frac{1\ lb}{453.6\ g} = \frac{26\ mL}{g}$

(3×10^1 mL/L for 1 sig fig)

Estimate: $50/4 = 12; 1000/500 = 2; (12 \times 2)/1 \approx 24$ mL/g

1.50 (a) $1257\ mi \times \frac{1\ km}{0.62137\ mi} \times \frac{charge}{225\ km} = 8.99\ charges$

Since charges are integral events, 9 total charges are required. The trip begins with a full charge, so 8 additional charges during the trip are needed.

(b) $\frac{14\ m}{s} \times \frac{1\ km}{1 \times 10^3\ m} \times \frac{1\ mi}{1.6093\ km} \times \frac{60\ s}{1\ min} \times \frac{60\ min}{1\ hr} = 31\ mi/hr$

(c) $450\ in^3 \times \frac{(2.54)^3\ cm^3}{1\ in^3} \times \frac{1\ mL}{1\ cm^3} \times \frac{1 \times 10^{-3}\ L}{1\ mL} = 7.37\ L$

(d) $2.4 \times 10^5\ barrels \times \frac{42\ gal}{1\ barrel} \times \frac{4\ qt}{1\ gal} \times \frac{1\ L}{1.057\ qt} = 3.8 \times 10^7\ L$

1.51 $14.5\ ft \times 16.5\ ft \times 8.0\ ft = 1914 = 1.9 \times 10^3\ ft^3$ (2 sig figs)

$$1914\ ft^3 \times \frac{(1\ yd)^3}{(3\ ft)^3} \times \frac{(1\ m)^3}{(1.0936)^3\ yd^3} \times \frac{10^3\ dm^3}{1\ m^3} \times \frac{1\ L}{1\ dm^3} \times \frac{1.19\ g}{L} \times \frac{1\ kg}{1000\ g} = 64.4985 = 64\ kg\ air$$

Estimate: $1900/27 \approx 60; (60 \times 1)/1 \approx 60$ kg

1.52 $11.0 \text{ ft} \times 11.5 \text{ ft} \times 20.5 \text{ ft} = 2593.25 = 2.59 \times 10^3 \text{ ft}^3$

$$2593.25 \text{ ft}^3 \times \frac{(1 \text{ yd})^3}{(3 \text{ ft})^3} \times \frac{(1 \text{ m})^3}{(1.0936 \text{ yd})^3} \times \frac{48 \, \mu\text{g CO}}{1 \text{ m}^3} \times \frac{1 \times 10^{-6} \text{ g}}{1 \, \mu\text{g}} = 3.5 \times 10^{-3} \text{ g CO}$$

1.53 Select a common unit for comparison, in this case the cm.

1 in ≈ 2.5 cm, 1 m = 100 cm

57 cm = 57 cm

14 in ≈ 35 cm

1.1 m = 110 cm

The order of length from shortest to longest is 14-in shoe < 57-cm string < 1.1-m pipe.

1.54 Select a common unit for comparison, in this case the kg.

1 kg > 2 lb, 1 L ≈ 1 qt

5 lb potatoes < 2.5 kg

5 kg sugar = 5 kg

1 gal = 4 qt ≈ 4 L; 1 mL H_2O = 1 g H_2O; 1 L = 1000 g, 4 L = 4000 g = 4 kg

The order of mass from lightest to heaviest is 5 lb potatoes < 1 gal water < 5 kg sugar.

1.55 Strategy: 1) Calculate volume of gold (Au) in cm^3 in the sheet

2) Mass = density × volume

3) Change g → troy oz and $

$$100 \text{ ft} \times 82 \text{ ft} \times \frac{(12)^2 \text{ in}^2}{1 \text{ ft}^2} \times 5 \times 10^{-6} \text{in} \times \frac{(2.54)^3 \text{ cm}^3}{1 \text{ in}^3} = 96.75 = 1 \times 10^2 \text{ cm}^3 \text{ Au}$$

$$96.75 \text{ cm}^3 \text{ Au} \times \frac{19.32 \text{ g}}{1 \text{ cm}^3} \times \frac{1 \text{ troy oz}}{31.1034768 \text{ g}} \times \frac{\$953}{\text{troy oz}} = \$57{,}272 = \$6 \times 10^4$$

(Strictly speaking, the datum 100 ft has 1 sig fig, so the result has 1 sig fig.)

1.56 A wire is a very long, thin cylinder of volume, $V = \pi r^2 h$, where h is the length of the wire and πr^2 is the cross-sectional area of the wire.

Strategy: 1) Calculate total volume of copper in cm^3 from mass and density

2) $h \text{ (length in cm)} = \frac{V}{\pi r^2}$

3) Change cm → ft

$$150 \text{ lb Cu} \times \frac{453.6 \text{ g}}{1 \text{ lb Cu}} \times \frac{1 \text{ cm}^3}{8.94 \text{ g}} = 7610.7 = 7.61 \times 10^3 \text{ cm}^3$$

$$r = d/2 = 7.50 \text{ mm} \times \frac{1 \text{ cm}}{10 \text{ mm}} \times \frac{1}{2} = 0.375 \text{ cm}$$

$$h = \frac{V}{\pi r^2} = \frac{7610.7\ \text{cm}^3}{\pi(0.375)^2\ \text{cm}^2} = 1.7227 \times 10^4 = 1.72 \times 10^4\ \text{cm}$$

$$1.7227 \times 10^4\ \text{cm} \times \frac{1\ \text{in}}{2.54\ \text{cm}} \times \frac{1\ \text{ft}}{12\ \text{in}} = 565\ \text{ft}$$

(too difficult to estimate)

Additional Exercises

1.57 (a) A gold coin is probably a *solid solution.* Pure gold (element 79) is too soft and too valuable to be used for coinage, so other metals are added. However, the simple term "gold coin" does not give a specific indication of the other metals in the mixture.

A cup of coffee is a *solution* if there are no suspended solids (coffee grounds). It is a heterogeneous mixture if there are grounds. If cream or sugar is added, the homogeneity of the mixture depends on how thoroughly the components are mixed.

A wood plank is a *heterogeneous mixture* of various cellulose components. The different domains in the mixture are visible as wood grain or knots.

(b) The ambiguity in each of these examples is that the name of the substance does not provide a complete description of the material. We must rely on mental images, and these vary from person to person.

1.58 (a) A *hypothesis* is a possible explanation for certain phenomena based on preliminary experimental data. A *theory* may be more general, and has a significant body of experimental evidence to support it; a theory has withstood the test of experimentation.

(b) A scientific *law* is a summary or statement of natural behavior; it tells how matter behaves. A *theory* is an explanation of natural behavior; it attempts to explain why matter behaves the way it does.

1.59 Any sample of vitamin C has the same relative amount of carbon and oxygen; the ratio of oxygen to carbon in the isolated sample is the same as the ratio in synthesized vitamin C.

$$\frac{2.00\ \text{g O}}{1.50\ \text{g C}} = \frac{x\ \text{g O}}{6.35\ \text{g C}};\quad x = \frac{(2.00\ \text{g O})(6.35\ \text{g C})}{1.50\ \text{g C}} = 8.47\ \text{g O}$$

This calculation assumes the law of constant composition.

1.60 (a) I. $(22.52 + 22.48 + 22.54)/3 = 22.51$

II. $(22.64 + 22.58 + 22.62)/3 = 22.61$

Based on the average, set I is more accurate. That is, it is closer to the true value of 22.52%.

(b) Average deviation = Σ | value − average | /3

I. | 22.52 − 22.51 | + |22.48 − 22.51 | + |22.54 − 22.51 | /3 = 0.02

II. | 22.64 − 22.61 | + |22.58 − 22.61 | + |22.62 − 22.61 | /3 = 0.02

The two sets display the same precision, even though set I is more accurate.

1.61 (a) Appropriate. The number 22,727,000 implies a precision of one part per thousand, or 0.1%. This is an appropriate level of precision for the accounting records of a company like Apple Computer.

(b) Appropriate. Rainfall data can be measured to a precision of at least one decimal place. Calculating annual rainfall and average annual rainfall involves addition, which dictates that significant figures are determined by the least number of decimal places in the data being summed.

(c) Appropriate. The percentage has three significant figures. In a population as large as the United States, the number of people named Brown can surely be counted by census data or otherwise to a precision of three significant figures.

(d) Inappropriate. Letter grades are posted at most to two decimal places and three significant figures (if plus and minus modifiers are quantified). The grade-point-average, obtained by addition and division, cannot have more decimal places or significant figures than the numbers being averaged.

1.62 (a) volume (b) area (c) volume (d) density

(e) time (f) length (g) temperature

1.63 (a) $\frac{m}{s^2}$ (b) $\frac{kg-m}{s^2}$ (c) $\frac{kg-m}{s^2} \times m = \frac{kg-m^2}{s^2}$

(d) $\frac{kg-m}{s^2} \times \frac{1}{m^2} = \frac{kg}{m-s^2}$ (e) $\frac{kg-m^2}{s^2} \times \frac{1}{s} = \frac{kg-m^2}{s^3}$

(f) $\frac{m}{s}$ (g) $kg \times \left(\frac{m}{s}\right)^2 = \frac{kg-m^2}{s^2}$

1.64 (a) $2.4 \times 10^5 \text{ mi} \times \frac{1.609 \text{ km}}{1 \text{ mi}} \times \frac{1000 \text{ m}}{1 \text{ km}} = 3.862 \times 10^8 = 3.9 \times 10^8 \text{ m}$

(b) $2.4 \times 10^5 \text{ mi} \times \frac{1.609 \text{ km}}{1 \text{ mi}} \times \frac{1 \text{ hr}}{350 \text{ km}} \times \frac{60 \text{ min}}{1 \text{ hr}} \times \frac{60 \text{ s}}{1 \text{ min}} = 4.0 \times 10^6 \text{ s}$

(c) $3.862 \times 10^8 \text{ m} \times 2 \times \frac{1 \text{ s}}{3.00 \times 10^8 \text{ m}} = 2.574 = 2.6 \text{ s}$

(d) $\frac{29.783 \text{ km}}{\text{s}} \times \frac{1 \text{ mi}}{1.6093 \text{ km}} \times \frac{60 \text{ s}}{1 \text{ min}} \times \frac{60 \text{ min}}{1 \text{ hr}} = 6.6624 \times 10^4 \text{ mi/hr}$

1.65 (a) $575 \text{ ft} \times \frac{12 \text{ in}}{1 \text{ ft}} \times \frac{2.54 \text{ cm}}{1 \text{ in}} \times \frac{10 \text{ mm}}{1 \text{ cm}} \times \frac{1 \text{ quarter}}{1.55 \text{ mm}} = 1.1307 \times 10^5 = 1.13 \times 10^5 \text{ quarters}$

(b) $1.1307 \times 10^5 \text{ quarters} \times \dfrac{5.67 \text{ g}}{1 \text{ quarter}} = 6.41 \times 10^5 \text{ g (641 kg)}$

(c) $1.1307 \times 10^5 \text{ quarters} \times \dfrac{1 \text{ dollar}}{4 \text{ quarters}} = \$28{,}268 = \$2.83 \times 10^4$

(d) $\$11{,}687{,}233{,}914{,}811.11 \times \dfrac{1 \text{ stack}}{\$28{,}268} = 4.13 \times 10^8 \text{ stacks}$

1.66 (a) $\dfrac{\$1950}{\text{acre-ft}} \times \dfrac{1 \text{ acre}}{4840 \text{ yd}^2} \times \dfrac{3 \text{ ft}}{1 \text{ yd}} \times \dfrac{(1.094 \text{ yd})^3}{(1 \text{ m})^3} \times \dfrac{(1 \text{ m})^3}{(10 \text{ dm})^3} \times \dfrac{(1 \text{ dm})^3}{1 \text{ L}} =$

$\$1.583 \times 10^{-3}$/L or 0.1583 ¢/L (0.158 ¢/L to 3 sig figs)

(b) $\dfrac{\$1950}{\text{acre-ft}} \times \dfrac{1 \text{ acre-ft}}{2 \text{ households-year}} \times \dfrac{1 \text{ year}}{365 \text{ days}} \times 1 \text{ household} = \dfrac{\$2.671}{\text{day}} = \dfrac{\$2.67}{\text{day}}$

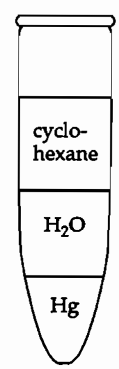
1.67 There are 347 degrees between the freezing and boiling points on the oleic acid (O) scale and 100 degrees on the celsius (C) scale. Also, 13°C = 0°O. By analogy with °F and °C,

$°\text{O} = \dfrac{100}{347}(°\text{C} - 13) \text{ or } °\text{C} = \dfrac{347}{100}(°\text{O}) + 13$

These equations correctly relate the freezing point (and boiling point) of oleic acid on the two scales.

f.p. of H_2O: $°\text{O} = \dfrac{100}{347}(0°\text{C} - 13) = -3.746 = -4°\text{O}$

1.68 The most dense liquid, Hg, will sink; the least dense, cyclohexane, will float; H_2O will be in the middle.



1.69 Density is the ratio of mass and volume. For samples with the same volume, in this case spheres with the same diameter, the denser ball will have a greater mass. The heavier ball, the red one on the right in the diagram is more dense.

1.70 The mass of water in the bottle does not change with temperature, but the density (ratio of mass to volume) does. That is, the amount of volume occupied by a certain mass of water changes with temperature. Calculate the mass of water in the bottle at 25°C, and then the volume occupied by this mass at –10°C.

(a) $25°\text{C}: 1.50 \text{ L } H_2O \times \dfrac{1000 \text{ cm}^3}{1 \text{ L}} \times \dfrac{0.997 \text{ g } H_2O}{1 \text{ cm}^3} = 1.4955 \times 10^3 = 1.50 \times 10^3 \text{ g } H_2O$

$-10°C$: $1.4955 \times 10^3 \text{ g } H_2O \times \frac{1 \text{ cm}^3}{0.917 \text{ g } H_2O} \times \frac{1 \text{ L}}{1000 \text{ cm}^3} = 1.6309 = 1.63 \text{ L}$

(b) No. If the soft-drink bottle is completely filled with 1.50 L of water, the 1.63 L of ice **cannot** be contained in the bottle. The extra volume of ice will push through any opening in the bottle, or crack the bottle to create an opening.

1.71 mass of toluene = 58.58 g − 32.65 g = 25.93 g

$\text{volume of toluene} = 25.93 \text{ g} \times \frac{1 \text{ mL}}{0.864 \text{ g}} = 30.0116 = 30.0 \text{ mL}$

volume of solid = 50.00 mL − 30.0116 mL = 19.9884 = 20.0 mL

$\text{density of solid} = \frac{32.65 \text{ g}}{19.9884 \text{ mL}} = 1.63 \text{ g/mL}$

1.72 (a) $V = 4/3\,\pi r^3 = 4/3\,\pi\,(28.9 \text{ cm})^3 = 1.0111 \times 10^5 = 1.01 \times 10^5 = 1.01 \times 10^5 \text{ cm}^3$

$1.0111 \times 10^5 \text{ cm}^3 \times \frac{19.3 \text{ g}}{\text{cm}^3} \times \frac{1 \text{ lb}}{453.59 \text{ g}} = 4{,}302 = 4.30 \times 10^3 \text{ lb}$

(b) No. The sphere weighs 4300 pounds, more than two tons. The student is unlikely to be able to carry the it without asssistance.

1.73 $1.00 \text{ gal battery acid} \times \frac{4 \text{ qt}}{1 \text{ gal}} \times \frac{1000 \text{ mL}}{1.0567 \text{ qt}} \times \frac{1.28 \text{ g}}{\text{mL}} = 4{,}845.3 = 4.85 \times 10^3 \text{ g battery acid}$

$4.8453 \times 10^3 \text{ g battery acid} \times \frac{38.1 \text{ g sulfuric acid}}{100 \text{ g battery acid}} = 1846 = 1.85 \times 10^3 \text{ g sulfuric acid}$

1.74 (a) $\frac{40 \text{ lb peat}}{14 \times 20 \times 30 \text{ in}^3} \times \frac{1 \text{ in}^3}{(2.54)^3 \text{ cm}^3} \times \frac{453.6 \text{ g}}{1 \text{ lb}} = 0.13 \text{ g/cm}^3 \text{ peat}$

$\frac{40 \text{ lb soil}}{1.9 \text{ gal}} \times \frac{1 \text{ gal}}{4 \text{ qt}} \times \frac{1.057 \text{ qt}}{1 \text{ L}} \times \frac{1 \times 10^{-3} \text{ L}}{1 \text{ mL}} \times \frac{1 \text{ mL}}{1 \text{ cm}^3} \times \frac{453.6 \text{ g}}{1 \text{ lb}} = 2.5 \text{ g/cm}^3 \text{ soil}$

No. Volume must be specified in order to compare mass. The densities tell us that a certain volume of peat moss is "lighter" (weighs less) than the same volume of top soil.

(b) $1 \text{ bag peat} = 14 \times 20 \times 30 = 8.4 \times 10^3 \text{ in}^3$

$15.0 \text{ ft} \times 20.0 \text{ ft} \times 3.0 \text{ in} \times \frac{12^2 \text{ in}^2}{\text{ft}^2} = 129{,}600 = 1.3 \times 10^5 \text{ in}^3 \text{ peat needed}$

$129{,}600 \text{ in}^3 \times \frac{1 \text{ bag}}{8.4 \times 10^3 \text{ in}^3} = 15.4 = 15 \text{ bags}$ (Buy 16 bags of peat.)

1.75 $8.0 \text{ oz} \times \frac{1 \text{ lb}}{16 \text{ oz}} \times \frac{453.6 \text{ g}}{\text{lb}} \times \frac{1 \text{ cm}^3}{2.70 \text{ g}} = 84.00 = 84 \text{ cm}^3$

$\frac{84 \text{ cm}^3}{50 \text{ ft}^2} \times \frac{1^2 \text{ ft}^2}{12^2 \text{ in}^2} \times \frac{1^2 \text{ in}^2}{2.54^2 \text{ cm}^2} \times \frac{10 \text{ mm}}{1 \text{ cm}} = 0.018 \text{ mm}$

1.76 $15\text{TW} \times \frac{1 \times 10^{12}\text{ W}}{1\text{TW}} \times \frac{1\text{ kW}}{1 \times 10^{3}\text{ W}} \times \frac{1^2\text{ m}^2}{1.336\text{ kW}} = 1.1228 \times 10^{10} = 1.1 \times 10^{10}\text{ m}^2$

Collection is 10% efficient, so 10 times this area is needed, $1.1 \times 10^{11}\text{ m}^2$.

$$1.97 \times 10^8\text{ mi}^2 \times \frac{1.6093^2\text{ km}^2}{1^2\text{ mi}^2} \times \frac{1000^2\text{ m}^2}{1^2\text{ km}^2} = 5.1020 \times 10^{14}\text{ m}^2 = 5.10 \times 10^{14}\text{ m}^2$$

$$\frac{1.1228 \times 10^{11}\text{ m}^2\text{ needed}}{5.1020 \times 10^{14}\text{ m}^2\text{ total surface Earth}} \times 100 = 0.02201 = 0.022\%\text{ Earth's surface needed}$$

1.77 $11.86\text{ g ethanol} \times \frac{1\text{ cm}^3}{0.789\text{ g ethanol}} = 15.0317 = 15.03\text{ cm}^3$, volume of cylinder

$$V = \pi r^2 h;\ r = (V/\pi h)^{1/2} = \left[\frac{15.0317\text{ cm}^3}{\pi \times 15.0\text{ cm}}\right]^{1/2} = 0.5648 = 0.565\text{ cm}$$

$d = 2r = 1.13\text{ cm}$

1.78 (a) Let x = mass of Au in jewelry

9.85 - x = mass of Ag in jewelry

The total volume of jewelry = volume of Au + volume of Ag

$$0.675\text{ cm}^3 = x\text{ g} \times \frac{1\text{ cm}^3}{19.3\text{ g}} + (9.85 - x)\text{g} \times \frac{1\text{ cm}^3}{10.5\text{ g}}$$

$$0.675 = \frac{x}{19.3} + \frac{9.85 - x}{10.5}$$ (To solve, multiply both sides by (19.3)(10.5))

$0.675\,(19.3)(10.5) = 10.5\,x + (9.85 - x)(19.3)$

$136.79 = 10.5\,x + 190.105 - 19.3\,x$

$-53.315 = -8.8\,x$

x = 6.06 g Au; 9.85 g total – 6.06 g Au = 3.79 g Ag

$$\text{mass \% Au} = \frac{6.06\text{ g Au}}{9.85\text{ g jewelry}} \times 100 = 61.5\%\text{ Au}$$

(b) 24 carats × 0.615 = 15 carat gold

1.79 The separation with distinctly separated red and blue spots is more successful. The procedure that produced the purple blur did not separate the two dyes. To quantify the characteristics of the separation, calculate a reference value for each spot that is

$$\frac{\text{distance travelled by spot}}{\text{distance travelled by solvent}}$$

If the values for the two spots are fairly different, the separation is successful. (One could measure the distance between the spots, but this would depend on the length of paper used and be different for each experiment. The values suggested above are independent of the length of paper.)

1.80 The densities are:

carbon tetrachloride (methane, tetrachloro) – 1.5940 g/cm^3

hexane – 0.6603 g/cm^3

benzene – 0.87654 g/cm^3

methylene iodide (methane, diiodo) – 3.3254 g/cm^3

Only methylene iodide will separate the two granular solids. The undesirable solid (2.04 g/cm^3) is less dense than methylene iodide and will float; the desired material is more dense than methylene iodide and will sink. The other three liquids are less dense than both solids and will not produce separation.

1.81 (a) $$10.0\text{ mg} \times \frac{1 \times 10^{-3}\text{ g}}{1\text{ mg}} \times \frac{1\text{ cm}^3}{0.20\text{ g}} = 0.050\text{ cm}^3 = 0.050\text{ mL volume}$$

(b) $$10.0\text{ mg} \times \frac{1 \times 10^{-3}\text{ g}}{1\text{ mg}} \times \frac{1242\text{ m}^2}{1\text{ g}} = 12.42 = 12.4\text{ m}^2\text{ surface area}$$

(c) 7.748 mg Hg initial – 0.001 mg Hg remain = 7.747 mg Hg removed

$$\frac{7.747\text{ mg Hg removed}}{7.748\text{ mg Hg initial}} \times 100 = 99.99\%\text{ Hg removed}$$

(d) 10.0 mg "spongy" initial + 7.747 mg Hg removed=17.747=17.7 mg after exposure

1.82 Study (a) is likely to be both precise and accurate, because the errors are carefully controlled. The secondary weight standard will be resistant to chemical and physical changes, the balance is carefully calibrated, and weighings are likely to be made by the same person. The relatively large number of measurements is likely to minimize the effect of random errors on the average value. The accuracy and precision of study (b) depend on the veracity of the participants' responses, which cannot be carefully controlled. It also depends on the definition of "comparable lifestyle." The percentages are not precise, because the broad definition of lifestyle leads to a range of results (scatter). The relatively large number of participants improves the precision and accuracy. In general, controlling errors and maximizing the number of data points in a study improves precision and accuracy.

2 Atoms, Molecules, and Ions

Visualizing Concepts

2.1 (a) The path of the charged particle bends because the particle is repelled by the negatively charged plate and attracted to the positively charged plate.

(b) Like charges repel and opposite charges attract, so the sign of the electrical charge on the particle is negative.

(c) The greater the magnitude of the charges, the greater the electrostatic repulsion or attraction. As the charge on the plates is increased, the bending will increase.

(d) As the mass of the particle increases and speed stays the same, linear momentum (mv) of the particle increases and bending decreases. (See **A Closer Look**: The Mass Spectrometer.)

2.2 (a) $$\%\text{ abundance} = \frac{\#\text{ of mass number} \times \text{particles}}{\text{total number of particles}} \times 100$$

12 red ^{293}Nv particles

8 blue ^{295}Nv particles

20 total particles

$$\%\text{ abundance } ^{293}\text{Nv} = \frac{12}{20} \times 100 = 60\%$$

$$\%\text{ abundance } ^{295}\text{Nv} = \frac{8}{20} \times 100 = 40\%$$

(b) Atomic weight (AW) is the same as average atomic mass.

Atomic weight (average atomic mass) = Σ fractional abundance $\times$ mass of isotope

AW of Nv = 0.60(293.15) + 0.40(295.15) = 293.95 amu

(Since % abundance was calculated by counting exact numbers of particles, assume % abundance is an exact number. Then, the number of significant figures in the AW is determined by the number of sig figs in the masses of the isotopes.)

2.3 In general, metals occupy the left side of the chart, and nonmetals the right side.

metals: red and green *nonmetals*: blue and yellow

alkaline earth metal: red *noble gas*: yellow

2.4 Since the number of electrons (negatively charged particles) does not equal the number of protons (positively charged particles), the particle is an ion. The charge on the ion is 2–.

Atomic number = number of protons = 16. The element is S, sulfur.

Mass number = protons + neutrons = 32

$^{32}_{16}S^{2-}$

2.5 In a solid, particles are close together and their relative positions are fixed. In a liquid, particles are close but moving relative to each other. In a gas, particles are far apart and moving. All ionic compounds are solids because of the strong forces among charged particles. Molecular compounds can exist in any state: solid, liquid, or gas.

Since the molecules in *ii* are far apart, *ii* must be a molecular compound. The particles in *i* are near each other and exist in a regular, ordered arrangement, so *i* is likely to be an ionic compound.

2.6 Formula: IF_5 Name: iodine pentafluoride
Since the compound is composed of elements that are all nonmetals, it is molecular.

2.7 See Figure 2.20.. yellow box: 1+ (group 1A); blue box: 2+ (group 2A)

black box: 3+ (a metal in Group 3A); red box: 2– (a nonmetal in group 6A);

green box: 1– (a nonmetal in group 7A)

2.8 Cations (red spheres) have positive charges; anions (blue spheres) have negative charges. There are twice as many anions as cations, so the formula has the general form CA_2. Only $Ca(NO_3)_2$, calcium nitrate, is consistent with the diagram.

Atomic Theory and the Discovery of Atomic Structure (sections 2.1-2.2)

2.9 Postulate 4 of the atomic theory is the *law of constant composition*. It states that the relative number and kinds of atoms in a compound are constant, regardless of the source. Therefore, 1.0 g of pure water should always contain the same relative amounts of hydrogen and oxygen, no matter where or how the sample is obtained.

2.10 (a) 6.500 g compound – 0.384 g hydrogen = 6.116 g sulfur

(b) *Conservation of mass*

(c) According to postulate 3 of the atomic theory, atoms are neither created nor destroyed during a chemical reaction. If 0.384 g of H are recovered from a compound that contains only H and S, the remaining mass must be sulfur.