Chapter Two: Atoms, Molecules, and Ions

Learning objectives

1. Identify all three hypotheses associated with Dalton's atomic theory and explain how they relate to the structure of matter and its associated reactions.
2. Recognize the importance of experiments conducted by Thomson, Millikan, Röntgen and Rutherford in regard to understanding the nature and structure of atoms.
3. Understand the different types of radiation that radioactive substances can produce.
4. Identify the location and physical properties of electrons, protons and neutrons in atoms.
5. Understand the nature and importance of isotopes.
6. Calculate the mass number of an isotope.
7. Utilize the mass number of an isotope to determine the number of electrons, protons or neutrons given other relevant information.
8. Utilize the periodic table to identify the chemical and physical properties of an element.
9. Categorize an element according to group based upon location in the periodic table.
10. Understand the nature of the atomic mass scale.
11. Calculate the average atomic mass of an element given the atomic mass and relative abundance of each of its naturally occurring isotopes.
12. Understand the information that chemical, molecular, and structural formulas provide.
13. Understand the differences between covalent and ionic bonding.
14. Determine the empirical formula of a compound given its molecular formula.
15. Utilize rules of nomenclature to name the different types of compounds including: covalent compounds, ionic compounds, oxoacids and hydrates.
16. Predict the charge of an ion formed from a main group element.
17. Name polyatomic ions and their associated charge.

Applications, Demonstrations, Tips and References

1. Page 35. Biological application.
2. Page 35. Instructor's Tip: If dimensional analysis did not identify the mathematically challenged students, calculating average atomic mass in this chapter will. Pointing out locations on campus where students can get help with their math skills may be helpful.
3. Page 35. Instructor's Tip: Historical footnotes can be used to bring meaning to some of the terms used in this chapter.
4. Page 36. Instructor's Tip: Nice lecture opener: Wouldn't it have been nice to study chemistry in ancient Greece? They only had four elements (earth, wind, water and fire).
5. Page 36. Literature: Mierrzecki, Roman."Dalton's atoms or Dalton's molecules? (SBS). *J. Chem. Educ.* **1981**, *58*, 1006.
6. Page 37. Organic application.
7. Page 37. Demonstration: You can easily demonstrate the law of conservation of mass using dry ice, an Erlenmeyer flask, scale and a balloon. The dry ice will sublime but no change in mass will occur.
8. Page 37. Multimedia: Law of conservation of mass.
9. Page 39. Instructor's Tip: Figures can be very helpful for illustrating the concept of an atom.

10. Page 39. Literature: Peake, Barrie M. "The discovery of the electron, proton, and neutron." *J. Chem. Educ.* **1989**, *66*, 738.
11. Page 39. Instructor's Tip: The word electron comes from the Greek word for amber. The term was introduced in 1891 by George Stoney.
12. Page 39. Multimedia: Cathode Ray Tube experiment
13. Page 40. Engineering application.
14. Page 40. Instructor's Tip: Consider talking about the production of various radioactive isotopes (iodine, cesium, etc.) that would be produced in a nuclear blast. This can be used to introduce the concept of isotopes and the fact that they just differ by the number of neutrons.
15. Page 40. Engineering application.
16. Page 41. Multimedia: Millikan Oil drop experiment
17. Page 42. Engineering application.
18. Page 43. Literature: Garrett, A. B. "The flash of genius 11. The neutron identified: Sir James Chadwick." *J. Chem. Educ.* **1962**, *39*, 638.
19. Page 44. Literature: Jensen, William B. "The Origins of the Symbols *A* and *Z* for Atomic Weight and Number." *J. Chem. Educ.* **2005**, *82*, 1764.
20. Page 44. Demonstration: Ellis, Arthur B. "Dramatizing isotopes: Deuterated ice cubes sink." *J. Chem. Educ.* **1990**, *67*, 159.
21. Page 44. Literature: Sein, Lawrence T., Jr. "Using Punnett Squares To Facilitate Students' Understanding of Isotopic Distributions in Mass Spectrometry." *J. Chem. Educ.* **2006**, *83*, 228.
22. Page 44. Engineering application.
23. Page 45. Literature: Laing, Michael. "The periodic table a new arrangement (PO)." *J. Chem. Educ.* **1989**, *66*, 746.
24. Page 46. Demonstration: Rizzo, Michelle M. et al. "Revisiting the Electric Pickle Demonstration." *J. Chem. Educ.* **2005**, *82*, 545.
25. Page 46. Literature: Hawkes, Stephen J. "Semimetallicity?" *J. Chem. Educ.* **2001**, *78*, 1686.
26. Page 46. Demonstration: Geselbracht, Margaret J. et al. "Mechanical Properties of Metals: Experiments with Steel, Copper, Tin, Zinc, and Soap Bubbles." *J. Chem. Educ.* **1994**, *71*, 254.
27. Page 47. Literature: Jensen, William B. "Why Helium Ends in "-ium."" *J. Chem. Educ.* **2004**, *81*, 944.
28. Page 47. Biological and environmental application.
29. Page 48. Engineering application.
30. Page 49. Engineering application.
31. Page 49. Environmental application.
32. Page 50. Literature: Last, Arthur M.; Webb, Michael J. "Using monetary analogies to teach average atomic mass (AA)." *J. Chem. Educ.* **1993**, *70*, 234.
33. Page 50. Literature: Schmid, Roland. "The Noble Gas Configuration --Not the Driving Force but the Rule of the Game in Chemistry." *J. Chem. Educ.* **2003**, *80*, 931.
34. Page 51. Literature: Jensen, William B. "The Origin of the Term Allotrope." *J. Chem. Educ.* **2006**, *83*, 838.
35. Page 51. Organic application.

36. Page 52. Instructor's Tip: Constructing a nomenclature flowchart can help students learn how to name compounds.
37. Page 52. Literature: Chimeno, Joseph. "How to Make Learning Chemical Nomenclature Fun, Exciting, and Palatable." *J. Chem. Educ.* **2000**, *77*, 144.
38. Page 54. Instructor's Tip: You can ask students to name some common acids and bases to stimulate their interest.
39. Page 54. Literature: Byrd, Shannon. "Learning the Functional Groups: Keys to Success." *J. Chem. Educ.* **2001**, *78*, 1355.
40. Page 54. Organic application.
41. Page 55. Organic application.
42. Page 56. Organic application.
43. Page 59. Instructor's Tip: It may help to stress why this system is not necessary for Group IA, IIA, IIIA metals (charge is predictable).
44. Page 60. Instructor's Tip: It is not necessary to make your students memorize all of the polyatomic ions, but you should clearly indicate which ones they will be responsible for.
45. Page 60. Multimedia: Electron transfer between sodium and chlorine
46. Page 62. Literature: Meek, Terry L. "Acidities of oxoacids: Correlation with charge distribution." *J. Chem. Educ.* **1992**, *69*, 270.
47. Page 62. Instructor's Tip: If students learn the rules for naming oxoacids, it makes it easier to remember some of the polyatomic ions (e.g. sulfite and sulfate).
48. Page 64. Literature: Schaeffer, Richard W et al. "Preparation and Analysis of Multiple Hydrates of Simple Salts." *J. Chem. Educ.* **2000**, *77*, 509.
49. Page 67. Biological application.

End of Chapter Problems sorted by difficulty

Easy
1, 2, 3, 4, 7, 8, 9, 10, 11, 12, 13, 16, 17, 18, 19, 20, 21, 22, 23, 24, 25, 26, 27, 28, 29, 30, 31, 32, 33, 34, 35, 37, 38, 40, 41, 42, 43, 48, 49, 50, 51, 52, 53, 54, 55, 56, 57, 58, 59, 60, 61, 62, 63, 64, 65, 66, 71, 72, 73, 74, 77, 78, 79, 80, 85, 86, 87, 88, 93, 97, 100, 101, 102, 108, 109

Medium
5, 6, 14, 15, 39, 44, 45, 46, 47, 67, 68, 69, 70, 75, 76, 81, 82, 83, 84, 89, 90, 91, 92, 94, 95, 96, 98, 99, 103, 104, 105, 106, 107, 110, 111, 112, 114, 115, 116, 117, 119, 120, 121, 122, 123, 125, 127

Difficult
36, 113, 118, 124, 126, 128

End of Chapter Problems sorted by type

Review
1, 2, 3, 4, 6, 5, 7, 8, 9, 10, 11, 12, 13, 16, 17, 18, 19, 29, 30, 31, 32, 33, 40, 41, 42, 43, 50, 51, 52, 53, 54, 55, 56, 57, 58, 71, 72, 73, 74

Conceptual
4, 6, 20, 22, 24, 26, 60, 62, 64, 88, 94, 113, 115, 116, 122

Biological
28, 33, 39, 65

Engineering
14, 33, 36, 119, 120, 124, 128

Environmental
None

Organic
65, 66, 104, 121, 123

Chapter 2

Atoms, Molecules, and Ions

Practice Problems C

2.1 (ii) and (iii)

2.2 (i) 14N, (ii) 21Na, (iii) 15O

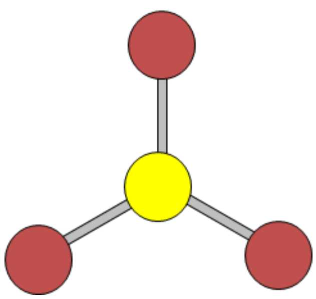
2.3 (i) 2.9177 g

(ii) 3.4679 g

(iii) 3.4988 g

2.4 4 ethanol molecules; all C used, 1 O left over, 3 H left over

2.5 selenium hexachloride

2.6 

2.7 formaldehyde and glucose

2.8 Fe(NO3)2, iron(II) nitrate

2.9 CuSO3, copper(II) sulfite

2.10 (i) and (iv)

2.11 (ii) and (iv)

Questions and Problems

2.1 1. Elements are composed of extremely small particles called atoms. All atoms of a given element are identical, having the same size, mass, and chemical properties. The atoms of one element are different from the atoms of all other elements.

2. Compounds are composed of atoms of more than one element. In any given compound, the same types of atoms are always present in the same relative numbers.

3. A chemical reaction rearranges atoms in chemical compounds; it does not create or destroy them.

2.2 The law of definite proportions states that different samples of a given compound always contain the same elements in the same mass ratio.

The law of multiple proportions states that if two elements can combine to form more than one compound with each other, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers.

2.3
$$\frac{\text{ratio of N to O in compound 1}}{\text{ratio of N to O in compound 2}} = \frac{0.8756}{0.4378} \approx \mathbf{2:1}$$

2.4
$$\frac{\text{ratio of P to Cl in compound 1}}{\text{ratio of P to Cl in compound 2}} = \frac{0.2912}{0.1747} \approx \mathbf{1.667} \approx \mathbf{5:3}$$

2.5
$$\frac{\text{ratio of F to S in } S_2F_{10}}{\text{ratio of F to S in } SF_4} = \frac{2.962}{2.370} \approx 1.250$$

$$\frac{\text{ratio of F to S in } SF_6}{\text{ratio of F to S in } SF_4} = \frac{3.555}{2.370} \approx 1.5$$

$$\frac{\text{ratio of F to S in } SF_4}{\text{ratio of F to S in } SF_4} = 1$$

$$\text{ratio in } SF_6 : \text{ratio in } S_2F_{10} : \text{ratio in } SF_4 = 1.5:1.25:1$$

Multiply through to get all whole numbers. $4\cdot(1.5:1.25:1) = \mathbf{6:5:4}$

2.6
$$\frac{\text{ratio of O to Fe in FeO}}{\text{ratio of O to Fe in } Fe_2O_3} = \frac{0.2865}{0.4297} \approx \mathbf{0.667} \approx \mathbf{2:3}$$

2.7
$$\frac{\text{g blue: 1.00 g red (right)}}{\text{g blue: 1.00 g red (left)}} = \frac{2/3}{1/1} \approx \mathbf{0.667} \approx \mathbf{2:3}$$

2.8
$$\frac{\text{g green: 1.00 g orange (right)}}{\text{g green: 1.00 g orange (left)}} = \frac{4/2}{3/1} \approx \mathbf{0.667:1} \approx \mathbf{2:3}$$

2.9 a. An α particle is a positively charged particle consisting of two protons and two neutrons, emitted in radioactive decay or nuclear fission.

b. A β particle is a high-speed electron, especially emitted in radioactive decay.

c. γ rays are high-energy electromagnetic radiation emitted by radioactive decay.

d. X-rays are a form of electromagnetic radiation similar to light but of shorter wavelength.

2.10 alpha rays, beta rays, and gamma rays

2.11 α particles are deflected away from positively charged plates. Cathode rays are drawn toward positively charged plates. Protons are positively charged particles in the nucleus. Neutrons are electrically neutral subatomic particles in the nucleus. Electrons are negatively charged particles that are distributed around the nucleus.

2.12 J .J. Thomson determined the ratio of electric charge to the mass of an individual electron.

R. A. Millikan calculated the mass of an individual electron and proved the charge on each electron was exactly the same.

Ernest Rutherford proposed that an atom's positive charges are concentrated in the nucleus and that most of the atom is empty space.

James Chadwick discovered neutrons.

2.13 Rutherford bombarded gold foil with α particles. Most of them passed through the foil, while a small proportion were deflected or reflected. Thus, most of the atom must be empty space through which the α particles could pass without encountering any obstructions.

2.14 First, convert 1 cm to picometers.

$$1\text{ cm} \times \frac{0.01\text{ m}}{1\text{ cm}} \times \frac{1\text{ pm}}{1 \times 10^{-12}\text{ m}} = 1 \times 10^{10}\text{ pm}$$

$$(1 \times 10^{10}\text{ pm}) \times \frac{1\text{ He atom}}{1 \times 10^{2}\text{ pm}} = \mathbf{1 \times 10^{8}\ He\ atoms}$$

2.15 Note that you are given information to set up the conversion factor relating meters and miles.

$$\boldsymbol{r}_{\textbf{atom}} = 10^4 r_{\text{nucleus}} = 10^4 \times 2.0\text{ cm} \times \frac{1\text{ m}}{100\text{ cm}} \times \frac{1\text{ mi}}{1609\text{ m}} = \mathbf{0.12\ mi}$$

2.16 Atomic number is the number of protons in the nucleus of each atom of an element. It determines the chemical identity of the element. There are 2 protons in each atom of helium-4.

Mass number is the total number of neutrons and protons present in the nucleus of an atom of an element. The mass number of helium-4 is 4. There are (4 – 2) = 2 neutrons in each atom.

Because atoms are electrically neutral, the number of protons and electrons must be equal. The atomic number is also the number of electrons in each atom.

2.17 The atomic number is the number of protons in the nucleus. It determines the chemical identity of the element. If an atom has a different number of protons (a different atomic number), it is a different element.

2.18 isotopes

2.19 *X* is the element symbol. It indicates the chemical identity of the atom.

A is the mass number. It is the number of protons plus the number of neutrons.

Z is the atomic number. It is the number of protons.

2.20 For iron, the atomic number *Z* is 26. Therefore the mass number *A* is:

$$\boldsymbol{A} = 26 + 28 = \mathbf{54}$$

2.21 **Strategy:** The 239 in Pu-239 is the mass number. The **mass number (*A*)** is the total number of neutrons and protons present in the nucleus of an atom of an element. You can look up the atomic number (number of protons) on the periodic table.

Setup:

Solution: mass number = number of protons + number of neutrons

number of neutrons = mass number – number of protons = 239 – 94 = **145**

2.22

Isotope	$^{3}_{2}He$	$^{4}_{2}He$	$^{24}_{12}Mg$	$^{25}_{12}Mg$	$^{48}_{22}Ti$	$^{79}_{35}Br$	$^{195}_{78}Pt$
No. Protons	2	2	12	12	22	35	78

No. Neutrons	1	2	12	13	26	44	117

2.23

Isotope	$^{15}_{7}N$	$^{33}_{16}S$	$^{63}_{29}Cu$	$^{84}_{38}Sr$	$^{130}_{56}Ba$	$^{186}_{74}W$	$^{202}_{80}Hg$
No. Protons	7	16	29	38	56	74	80
No. Electrons	7	16	29	38	56	74	80
No. Neutrons	8	17	34	46	74	112	122

2.24 a. $^{23}_{11}Na$ b. $^{64}_{28}Ni$ c. $^{115}_{50}Sn$ d. $^{42}_{20}Ca$

2.25 The accepted way to denote the atomic number and mass number of an element X is $^{A}_{Z}X$ where $\boldsymbol{A}$ = mass number and $\boldsymbol{Z}$ = atomic number.

a. $^{186}_{74}W$ b. $^{201}_{80}Hg$ c. $^{76}_{34}Se$ d. $^{239}_{94}Pu$

2.26 a. 10 b. 26 c. 81 d. 196

2.27 a. 19 b. 34 c. 75 d. 192

2.28 ^{198}Au: 119 neutrons, ^{47}Ca: 27 neutrons, ^{60}Co: 33 neutrons, ^{18}F: 9 neutrons, ^{125}I: 72 neutrons, ^{131}I: 78 neutrons, ^{42}K: 23 neutrons, ^{43}K: 24 neutrons, ^{24}Na: 13 neutrons, ^{32}P: 17 neutrons, ^{85}Sr: 47 neutrons, ^{99}Tc: 56 neutrons.

2.29 The periodic table is a chart in which elements having similar chemical and physical properties are grouped together.

2.30 A metal is a good conductor of heat and electricity, whereas a nonmetal is usually a poor conductor of heat and electricity.

2.31 Answers will vary.

2.32 Answers will vary.

2.33 Strontium has similar chemical properties to calcium, which is an important mineral for humans.

2.34 Helium and Selenium are nonmetals whose name ends with *ium*. (Tellurium is a metalloid whose name ends in *ium*.)

2.35 a. Metallic character increases as you progress down a group of the periodic table. For example, moving down Group 4A, the nonmetal carbon is at the top and the metal lead is at the bottom of the group.

b. Metallic character decreases from the left side of the table (where the metals are located) to the right side of the table (where the nonmetals are located).

2.36 a. Li (0.53 g/cm^3) K (0.86 g/cm^3) H_2O (0.98 g/cm^3)

b. Au (19.3 g/cm^3) Pt (21.4 g/cm^3) Hg (13.6 g/cm^3)

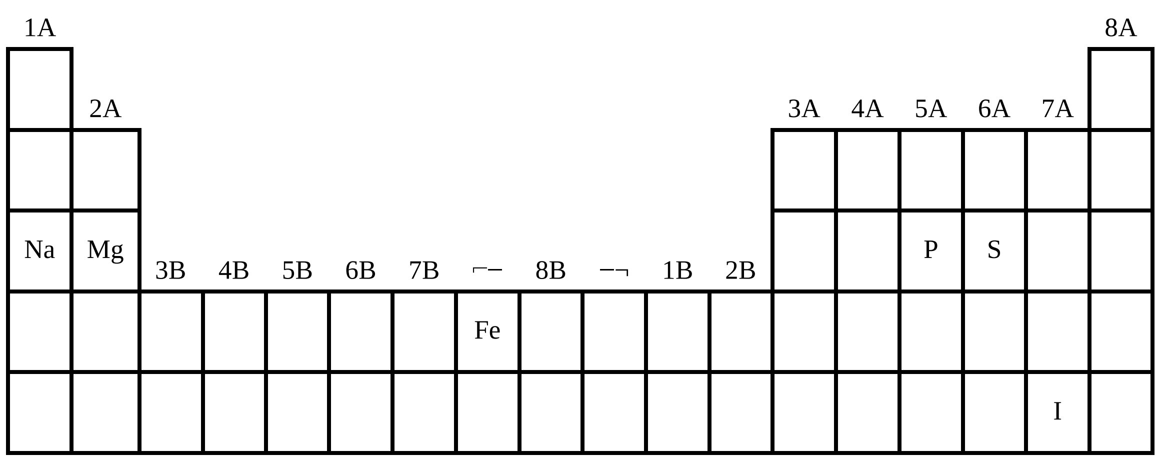
c. Os (22.6 g/cm^3)

d. Te (6.24 g/cm^3)

2.37 Na and K are both Group 1A elements; they should have similar chemical properties. N and P are both Group 5A elements; they should have similar chemical properties. F and Cl are Group 7A elements; they should have similar chemical properties.

2.38 I and Br (both in Group 7A), O and S (both in Group 6A), Ca and Ba (both in Group 2A)

2.39



Atomic number 26, iron, Fe, (present in hemoglobin for transporting oxygen)

Atomic number 53, iodine, I, (present in the thyroid gland)

Atomic number 11, sodium, Na, (present in intra- and extra-cellular fluids)

Atomic number 15, phosphorus, P, (present in bones and teeth)

Atomic number 16, sulfur, S, (present in proteins)

Atomic number 12, magnesium, Mg, (present in chlorophyll molecules)

2.40 An atomic mass unit (amu) is defined as a mass exactly equal to one-twelfth the mass of one carbon-12 atom. Setting the atomic mass of carbon-12 at 12 amu provides the standard for measuring the atomic mass of the other elements, since the mass of a single atom cannot be measured.

2.41 The mass of a carbon-12 atom is exactly 12 amu. The mass on the periodic table is the average mass of naturally occurring carbon, which is a mixture of several carbon isotopes.

2.42 The average mass of the naturally occurring isotopes of gold, taking into account their natural abundances, is 197.0 amu.

2.43 To calculate the average atomic mass of an element, you must know the identity and natural abundances of all naturally occurring isotopes of the element.

2.44

$$(34.968 \text{ amu})(0.7553) + (36.956 \text{ amu})(0.2447) = \mathbf{35.45\ amu}$$

2.45 $(203.973020 \text{ amu})(0.014) + (205.974440 \text{ amu})(0.241) + (206.975872 \text{ amu})(0.221) + (207.976627 \text{ amu})(0.524) = \mathbf{207.2\ amu}$

2.46 The fractional abundances of the two isotopes of Tl must add to 1. Therefore, we can write

$$(202.972320 \text{ amu})(x) + (204.974401 \text{ amu})(1 - x) = 204.4 \text{ amu}$$

Solving for x gives 0.2869. Therefore, the natural abundances of ^{203}Tl and ^{205}Tl are **28.69%** and **71.31%**, respectively.

2.47 **Strategy:** Each isotope contributes to the average atomic mass based on its relative abundance. Multiplying the mass of an isotope by its fractional abundance (not percent) will give the contribution to the average atomic mass of that particular isotope.

It would seem that there are two unknowns in this problem, the fractional abundance of ^{6}Li and the fractional abundance of ^{7}Li. However, these two quantities are not independent of each other; they are related by the fact that they must sum to 1. Start by letting *x* be the fractional abundance of ^{6}Li. Since the sum of the two fractional abundances must be 1, we can write

$$(6.0151\text{ amu})(x) + (7.0160\text{ amu})(1 - x) = 6.941\text{ amu}$$

Setup:

Solution: Solving for *x* gives 0.075, which corresponds to the fractional abundance of ^{6}Li. The fractional abundance of ^{7}Li is $(1 - x) = 0.925$. Therefore, the natural abundances of ^{6}Li and ^{7}Li are **7.5%** and **92.5%**, respectively.

2.48 The conversion factor required is $\dfrac{6.022\times10^{23}\text{ amu}}{1\text{ g}}$

$$13.2\text{ amu} \times \frac{1\text{ g}}{6.022\times10^{23}\text{ amu}} = \mathbf{2.19\times10^{-23}\text{ g}}$$

2.49 The conversion factor required is $\dfrac{6.022\times10^{23}\text{ amu}}{1\text{ g}}$

$$8.4\text{ g} \times \frac{6.022\times10^{23}\text{ amu}}{1\text{ g}} = \mathbf{5.1\times10^{24}\text{ amu}}$$

2.50 A molecule is a combination of at least two atoms in a specific arrangement held together by electrostatic forces known as covalent chemical bonds.

2.51 An allotrope is one of two or more distinct forms of an element. For example, diamond and graphite are two allotropes of carbon. Allotropes have different chemical bonding of atoms of the same element. Isotopes have different nuclear structures.

2.52 Two common molecular models are ball-and-stick and space-filling.

2.53 A chemical formula denotes the composition of the substance.

a. 1:1

b. 1:3

c. 2:4 = 1:2

d. 4:6 = 2:3

2.54 A molecular formula shows the exact number of atoms of each element in a molecule. An empirical formula shows the lowest whole number ratio of the atoms of each element in a molecule.

2.55 Answers will vary. Example: C_2H_4 and C_4H_8

2.56 Organic compounds contain carbon and hydrogen, sometimes in combination with other elements such as oxygen, nitrogen, sulfur, and the halogens. Inorganic compounds generally do not contain carbon, although some carbon-containing species are considered inorganic.

2.57 Answers will vary.

Binary: carbon dioxide, CO_2

Ternary: dichloromethane, CH_2Cl_2

2.58 HCl in the gas phase is hydrogen chloride, a molecular compound. When dissolved in water, it dissociates completely into ions and is hydrochloric acid.

2.59 a. This is a polyatomic molecule that is an elemental form of the substance. It is not a compound.

b. This is a polyatomic molecule that is a compound.

c. This is a diatomic molecule that is a compound.

2.60 a. This is a diatomic molecule that is a compound.

b. This is a polyatomic molecule that is a compound.

c. This is a polyatomic molecule that is the elemental form of the substance. It is not a compound.

2.61 **Elements:** N_2, S_8, H_2

Compounds: NH_3, NO, CO, CO_2, SO_2

2.62 There are more than two correct answers for each part of the problem.

a. H_2 and F_2

b. HCl and CO

c. S_8 and P_4

d. H_2O and $C_{12}H_{22}O_{11}$ (sucrose)

2.63 **Strategy:** An *empirical formula* tells us which elements are present and the *simplest* whole-number ratio of their atoms. Can you divide the subscripts in the formula by a common factor to end up with smaller whole-number subscripts?

Setup:

Solution: a. Dividing both subscripts by 2, the simplest whole number ratio of the atoms in C_2N_2 is **CN**.

b. Dividing all subscripts by 6, the simplest whole number ratio of the atoms in C_6H_6 is **CH**.

c. The molecular formula as written, $\mathbf{C_9H_{20}}$, contains the simplest whole number ratio of the atoms present. In this case, the molecular formula and the empirical formula are the same.

d. Dividing all subscripts by 2, the simplest whole number ratio of the atoms in P_4O_{10} is $\mathbf{P_2O_5}$.

e. Dividing all subscripts by 2, the simplest whole number ratio of the atoms in B_2H_6 is $\mathbf{BH_3}$.

Think About It:

2.64 a. $AlBr_3$ b. $NaSO_2$ c. N_2O_5 d. $K_2Cr_2O_7$ e. HCO_2

2.65 $C_3H_7NO_2$

2.66 C_2H_6O (The formula for ethanol can also be written as C_2H_5OH or CH_3CH_2OH.)

2.67 a. nitrogen trichloride

b. iodine heptafluoride

c. tetraphosphorus hexoxide

d. disulfur dichloride

2.68 a. PBr_3 b. N_2F_4 c. XeO_4 d. SeO_3

2.69 All of these are molecular compounds. We use prefixes to express the number of each atom in the molecule. The molecular formulas and names are:

a. NF_3: nitrogen trifluoride

b. PBr_5: phosphorus pentabromide

c. SCl_2: sulfur dichloride

2.70 a. OF_2: oxygen difluoride

b. Al_2Br_6: dialuminum hexabromide

c. N_2F_4: dinitrogen tetrafluoride (also "perfluorohydrazine")

2.71 Answers will vary.

2.72 An ionic compound consists of anions and cations. The ratio of anions and cations is such that the net charge is zero.

2.73 The formulas of ionic compounds are generally empirical formulas because an ionic compound consists of a vast array of interspersed cations and anions called a lattice, not discrete molecular units.

2.74 The Stock system uses Roman numerals to indicate the charge on cations of metals that commonly have more than one possible charge. This eliminates the need to know which charges are common on all the transition metals.

2.75 The **atomic number (*Z*)** is the number of protons in the nucleus of each atom of an element. You can find this on a periodic table. The number of **electrons** in an *ion* is equal to the number of protons minus the charge on the ion.

number of electrons (ion) = number of protons – charge on the ion

Ion	Na^{+}	Ca^{2+}	Al^{3+}	Fe^{2+}	I^{-}	F^{-}	S^{2-}	O^{2-}	N^{3-}
No. protons	11	20	13	26	53	9	16	8	7
No. electrons	10	18	10	24	54	10	18	10	10

2.76

Ion	K^{+}	Mg^{2+}	Fe^{3+}	Br^{-}	Mn^{2+}	C^{4-}	Cu^{2+}
No. protons	19	12	26	35	25	6	29
No. electrons	18	10	23	36	23	10	27

2.77 a. Sodium ion has a +1 charge and oxide has a –2 charge. The correct formula is **Na_2O**.

b. The iron ion has a +2 charge and sulfide has a –2 charge. The correct formula is **FeS**.

c. The correct formula is **$Co_2(SO_4)_3$**.

d. Barium ion has a +2 charge and fluoride has a –1 charge. The correct formula is **BaF_2**.

2.78 a. The copper ion has a +1 charge and bromide has a –1 charge. The correct formula is **CuBr**.

b. The manganese ion has a +3 charge and oxide has a –2 charge. The correct formula is **Mn_2O_3**.

c. We have the Hg_2^{2+} ion and iodide (I^-). The correct formula is **Hg_2I_2**.

d. Magnesium ion has a +2 charge and phosphate has a –3 charge. The correct formula is **$Mg_3(PO_4)_2$**.

2.79 Compounds of metals with nonmetals are usually ionic. Nonmetal-nonmetal compounds are usually molecular.

Ionic: LiF, $BaCl_2$, KCl

Molecular: $SiCl_4$, B_2H_6, C_2H_4

2.80 Compounds of metals with nonmetals are usually ionic. Nonmetal-nonmetal compounds are usually molecular.

Ionic: NaBr, BaF_2, CsCl.

Molecular: CH_4, CCl_4, ICl, NF_3

2.81 **Strategy:** When naming ionic compounds, our reference for the names of cations and anions are Tables 2.8 and 2.9 of the text. Keep in mind that if a metal can form cations of different charges, we need to use the Stock system. In the Stock system, Roman numerals are used to specify the charge of the cation. The metals that have only one charge in ionic compounds are the alkali metals (+1), the alkaline earth metals (+2), Ag^+, Zn^{2+}, Cd^{2+}, and Al^{3+}.

When naming acids, binary acids are named differently than oxoacids. For binary acids, the name is based on the nonmetal. For oxoacids, the name is based on the polyatomic anion. For more detail, see Section 2.7 of the text.

Setup:

Solution:

a. This is an ionic compound in which the metal cation (K^+) has only one charge. The correct name is **potassium dihydrogen phosphate**.

b. This is an ionic compound in which the metal cation (K^+) has only one charge. The correct name is **potassium hydrogen phosphate**

c. This is molecular compound. In the gas phase, the correct name is **hydrogen bromide**.

d. The correct name of this compound in water is **hydrobromic acid**.

e. This is an ionic compound in which the metal cation (Li^+) has only one charge. The correct name is **lithium carbonate**.

f. This is an ionic compound in which the metal cation (K^+) has only one charge. The correct name is **potassium dichromate**.

g. This is an ionic compound in which the cation is a polyatomic ion with a charge of +1. The anion is an oxoanion with one less O atom than the corresponding –ate ion (nitrate). The correct name is **ammonium nitrite**.

h. The oxoanion in this acid is analogous to the chlorate ion. The correct name of this compound is **hydrogen iodate (in water, iodic acid)**

i. This is a molecular compound. We use a prefix to denote how many F atoms it contains. The correct name is **phosphorus pentafluoride**.

j. This is a molecular compound. We use prefixes to denote the numbers of both types of atom. The correct name is **tetraphosphorus hexoxide**.

k. This is an ionic compound in which the metal cation (Cd^{2+}) has only one charge. The correct name is **cadmium iodide**.

l. This is an ionic compound in which the metal cation (Sr^{2+}) has only one charge. The correct name is **strontium sulfate**.

m. This is an ionic compound in which the metal cation (Al^{3+}) has only one charge. The correct name is **aluminum hydroxide**.

2.82
a. potassium hypochlorite
b. silver carbonate
c. nitrous acid
d. potassium permanganate
e. cesium chlorate
f. potassium ammonium sulfate
g. iron(II) oxide
h. iron(III) oxide
i. titanium(IV) chloride
j. sodium hydride
k. lithium nitride
l. sodium oxide
m. sodium peroxide

2.83 **Strategy:** When writing formulas of molecular compounds, the prefixes specify the number of each type of atom in the compound.

When writing formulas of ionic compounds, the subscript of the cation is numerically equal to the charge of the anion, and the subscript of the anion is numerically equal to the charge on the cation. If the charges of the cation and anion are numerically equal, then no subscripts are necessary. Charges of common cations and anions are listed in Tables 2.8 and 2.9 of the text. Keep in mind that Roman numerals specify the charge of the cation, *not* the number of metal atoms. Remember that a Roman numeral is not needed for some metal cations, because the charge is known. These metals are the alkali metals (+1), the alkaline earth metals (+2), Ag^+, Zn^{2+}, Cd^{2+}, and Al^{3+}.

When writing formulas of oxoacids, you must know the names and formulas of polyatomic anions (see Table 2.9 of the text).

Setup:

Solution:

a. Rubidium is an alkali metal. It only forms a +1 cation. The polyatomic ion nitrite, NO_2^-, has a –1 charge. Because the charges on the cation and anion are numerically equal, the ions combine in a one-to-one ratio. The correct formula is **$RbNO_2$**.

b. Potassium is an alkali metal. It only forms a +1 cation. The anion, sulfide, has a charge of –2. Because the charges on the cation and anion are numerically different, the subscript of the cation is numerically equal to the charge on the anion, and the subscript of the anion is numerically equal to the charge on the cation. The correct formula is **K_2S**.

c. Sodium is an alkali metal. It only forms a +1 cation. The anion is the *hydrogen sulfide* ion (the sulfide ion plus one hydrogen), HS^-. Because the charges are numerically the same, the ions combine in a one-to-one ratio.The correct formula is **NaHS**.

d. Magnesium is an alkaline earth metal. It only forms a +2 cation. The polyatomic phosphate anion has a charge of –3, PO_4^{3-}. Because the charges on the cation and anion are numerically different, the subscript of the cation is numerically equal to the charge on the anion, and the subscript of the anion is numerically equal to the charge on the cation. The correct formula is **$Mg_3(PO_4)_2$**. Note that for its subscript to be changed, a polyatomic ion must be enclosed in parentheses.

e. Calcium is an alkaline earth metal. It only forms a +2 cation. The polyatomic ion hydrogen phosphate, HPO_4^{2-}, has a –2 charge. Because the charges are numerically the same, the ions combine in a one-to-one ratio. The correct formula is **$CaHPO_4$**.

f. Lead (II), Pb^{2+}, is a cation with a charge +2. The polyatomic ion carbonate, CO_3^{2-}, has a –2 charge. Because the charges on the cation and anion are numerically equal, the ions combine in a one-to-one ratio. The correct formula is **$PbCO_3$**.

g. Tin (II), Sn^{2+}, is a cation with a charge of +2. The anion, fluoride, has a change of –1. Because the charges on the cation and anion are numerically different, the subscript of the cation is numerically equal to the charge on the anion, and the subscript of the anion is numerically equal to the charge on the cation. The correct formula is **SnF_2**.

h. The polyatomic ion ammonium, NH_4^+, has a +1 charge and the polyatomic ion sulfate, SO_4^{2-}, has a –2 charge. To balance the charge, we need 2 NH_4^+ cations. The correct formula is **$(NH_4)_2SO_4$**.

i. Silver forms only a +1 ion. The perchlorate ion, ClO_4^-, has a charge of –1. Because the charges are numerically the same, the ions combine in a one-to-one ratio. The correct formula is **$AgClO_4$**.

j. This is a molecular compound. The Greek prefixes tell you the number of each type of atom in the molecule: no prefix indicates 1 and tri- indicates 3. The correct formula is **BCl_3**.

2.84 a. $CuCN$ d. $HI(aq)$ g. IF_7 j. Hg_2I_2

b. $Sr(ClO_2)_2$ e. $Na_2(NH_4)PO_4$ h. P_4S_{10} k. SeF_6

c. $HBrO_4(aq)$ f. KH_2PO_4 i. HgO

2.85 a. $Mg(NO_3)_2$ b. Al_2O_3 b. LiH b. Na_2S

2.86 a. one green sphere, one red sphere

b. one green sphere, two red spheres

c. three green spheres, two red spheres

d. two green spheres, one red sphere

2.87 acid: compound that produces H^+; base: compound that produces OH^-; oxoacids: acids that contain oxygen; oxoanions: the anions that remain when oxoacids lose H^+ ions; hydrates: ionic solids that have water molecules in their formulas.

2.88 Uranium is radioactive. It loses mass because it constantly emits alpha (α) particles.

2.89 **(c)** Changing the electrical charge of an atom usually has a major effect on its chemical properties. The two electrically neutral carbon isotopes should have nearly identical chemical properties.

2.90 The number of protons = 65 − 35 = 30. The element that contains 30 protons is zinc, Zn. There are two fewer electrons than protons, so the charge of the cation is +2. The symbol for this cation is $\mathbf{Zn^{2+}}$.

2.91 Atomic number = 127 − 74 = 53. This anion has 53 protons, so it is an iodide ion. Since there is one more electron than protons, the ion has a −1 charge. The correct symbol is $\mathbf{I^-}$.

2.92 a. Species with the same number of protons and electrons will be neutral. **A, F, G**.

b. Species with more electrons than protons will have a negative charge. **B, E**.

c. Species with more protons than electrons will have a positive charge. **C, D**.

d. **A**: ${}^{10}_{5}\text{B}$ **B**: ${}^{14}_{7}\text{N}^{3-}$ **C**: ${}^{39}_{19}\text{K}^{+}$ **D**: ${}^{66}_{30}\text{Zn}^{2+}$ **E**: ${}^{81}_{35}\text{Br}^{-}$ **F**: ${}^{11}_{5}\text{B}$ **G**: ${}^{19}_{9}\text{F}$

2.93 NaCl is an ionic compound; it doesn’t consist of molecules.

2.94 **Yes**. The law of multiple proportions requires that the masses of sulfur combining with phosphorus must be in the ratios of small whole numbers. For the three compounds shown, four phosphorus atoms combine with three, seven, and ten sulfur atoms, respectively. If the atom ratios are in small whole number ratios, then the mass ratios must also be in small whole number ratios.

2.95 The species and their identification are as follows:

a. SO_2	molecule and compound	g. O_3	element and molecule

b. S_8	element and molecule	h. CH_4	molecule and compound
c. Cs	element	i. KBr	compound, not molecule
d. N_2O_5	molecule and compound	j. S	element
e. O	element	k. P_4	element and molecule
f. O_2	element and molecule	l. LiF	compound, not molecule

2.96 a. This is an ionic compound. Prefixes are *not* used. The correct name is barium chloride.

b. Iron has a +3 charge in this compound. The correct name is iron(III) oxide.

c. NO_2^- is the nitrite ion. The correct name is cesium nitrite.

d. Magnesium is an alkaline earth metal, which always has a +2 charge in ionic compounds. The roman numeral is not necessary. The correct name is magnesium bicarbonate.

2.97 All masses are relative, which means that the mass of every object is compared to the mass of a standard object (such as the piece of metal in Paris called the "standard kilogram"). The mass of the standard object is determined by an international committee, and that mass is an arbitrary number to which everyone in the scientific community agrees.

Atoms are so small it is hard to compare their masses to the standard kilogram. Instead, we compare atomic masses to the mass of one specific atom. In the 19th century the atom was ^{1}H, and for a good part of the 20th century it was ^{16}O. Now it is ^{12}C, which is given the arbitrary mass of 12 amu exactly. All other isotopic masses (and therefore average atomic masses) are measured relative to the assigned mass of ^{12}C.

2.98 a. Ammonium is NH_4^+, not NH_3^+. The formula should be $\mathbf{(NH_4)_2CO_3}$.

b. Calcium has a +2 charge and hydroxide has a −1 charge. The formula should be $\mathbf{Ca(OH)_2}$.

c. Sulfide is S^{2-}, not SO_3^{2-}. The correct formula is **CdS**.

d. Dichromate is $Cr_2O_7^{2-}$, not $Cr_2O_4^{2-}$. The correct formula is $\mathbf{ZnCr_2O_7}$.

2.99

Symbol	$^{11}_{5}B$	$^{54}_{26}Fe^{2+}$	$^{31}_{15}P^{3-}$	$^{196}_{79}Au$	$^{222}_{86}Rn$
Protons	5	26	15	79	86
Neutrons	6	28	16	117	136
Electrons	5	24	18	79	86
Net Charge	0	+2	–3	0	0

2.100 a. Ionic compounds are typically formed between metallic and nonmetallic elements.

b. In general the transition metals, the actinides and lanthanides have variable charges.

2.101 a. Li^+, alkali metals always have a +1 charge in ionic compounds

b. S^{2-}

c. I^-, halogens have a –1 charge in ionic compounds

d. N^{3-}

e. Al^{3+}, aluminum always has a +3 charge in ionic compounds

f. Cs^+, alkali metals always have a +1 charge in ionic compounds

g. Mg^{2+}, alkaline earth metals always have a +2 charge in ionic compounds.

2.102 The symbol ^{23}Na provides more information than $_{11}Na$. The mass number plus the chemical symbol identifies a specific isotope of Na (sodium) while combining the atomic number with the chemical symbol tells you nothing new. Can other isotopes of sodium have different atomic numbers?

2.103 The binary Group 7A element acids are: HF, hydrofluoric acid; HCl, hydrochloric acid; HBr, hydrobromic acid; HI, hydroiodic acid. Oxoacids containing Group 7A elements (using the specific examples for chlorine) are: $HClO_4$, perchloric acid; $HClO_3$, chloric acid; $HClO_2$, chlorous acid: HClO, hypochlorous acid.

Examples of oxoacids containing other Group A-block elements are: H_3BO_3, boric acid (Group 3A); H_2CO_3, carbonic acid (Group 4A); HNO_3, nitric acid and H_3PO_4, phosphoric acid (Group 5A); and H_2SO_4, sulfuric acid (Group 6A). Hydrosulfuric acid, H_2S, is an example of a binary Group 6A acid while HCN, hydrocyanic acid, contains both a Group 4A and 5A element.

2.104 a. C_2H_2, CH a. C_6H_6, CH a. C_2H_6, CH_3 a. C_3H_8, C_3H_8

2.105 a.

Isotope	$^{4}_{2}He$	$^{20}_{10}Ne$	$^{40}_{18}Ar$	$^{84}_{36}Kr$	$^{132}_{54}Xe$
No. Protons	2	10	18	36	54
No. Neutrons	2	10	22	48	78

b.

neutron/proton ratio	1.00	1.00	1.22	1.33	1.44

The neutron/proton ratio increases with increasing atomic number.

2.106 H_2, N_2, O_2, F_2, Cl_2, He, Ne, Ar, Kr, Xe, Rn

2.107 Cu, Ag, and Au are fairly chemically unreactive. This makes them especially suitable for making coins and jewelry that you want to last a very long time.

2.108 They generally do not react with other elements. Helium, neon, and argon are chemically inert.

2.109 Magnesium and strontium are also alkaline earth metals. You should expect the charge of the metal to be the same (+2). **MgO** and **SrO**.

2.110 All isotopes of radium are radioactive. It is a radioactive decay product of uranium-238. Radium itself does *not* occur naturally on Earth.

2.111 a. $\frac{2\text{ red} : 1\text{ blue}}{1\text{ red} : 1\text{ blue}} = \mathbf{2:1}$

b. $\frac{1\text{ red} : 2\text{ blue}}{1\text{ red} : 1\text{ blue}} = \mathbf{1:2}$

c. $\dfrac{4 \text{ red} : 2 \text{ blue}}{1 \text{ red} : 1 \text{ blue}} = 4:2 = \mathbf{2:1}$

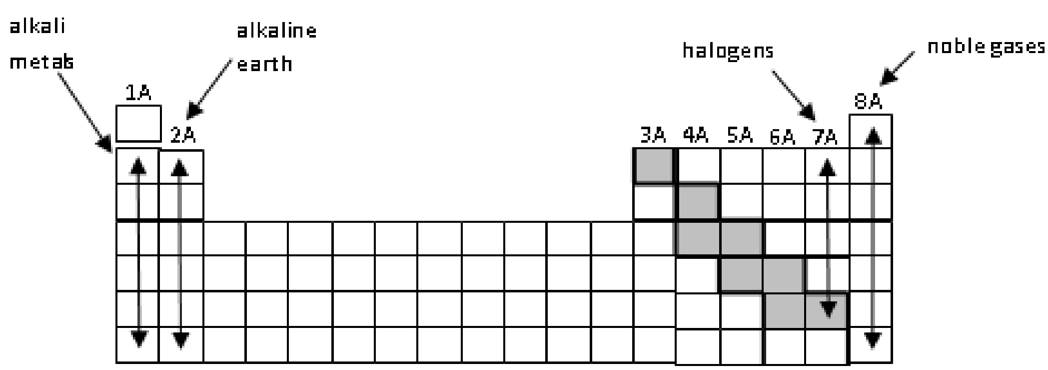
d. $\dfrac{5 \text{ red} : 2 \text{ blue}}{1 \text{ red} : 1 \text{ blue}} = \mathbf{5:2}$

2.113 The mass of fluorine reacting with hydrogen and deuterium would be the same. The ratio of F atoms to hydrogen (or deuterium) atoms is 1:1 in both compounds. This does not violate the law of definite proportions. When the law of definite proportions was formulated, scientists did not know of the existence of isotopes.

2.114 a. NaH, sodium hydride
c. Na_2S, sodium sulfide
e. OF_2, oxygen difluoride
b. B_2O_3, diboron trioxide
d. AlF_3, aluminum fluoride
f. $SrCl_2$, strontium chloride

2.115 a. Br a. Rn a. Se a. Rb a. Pb

2.116



The metalloids are shown in gray.

2.117

Cation	Anion	Formula	Name
Mg^{2+}	**HCO_3^-**	**$Mg(HCO_3)_2$**	Magnesium bicarbonate
Sr^{2+}	**Cl^-**	$SrCl_2$	**Strontium chloride**

Fe^{3+}	NO_2^-	**$Fe(NO_2)_3$**	**Iron(III) nitrite**
Mn^{2+}	**ClO_3^-**	**$Mn(ClO_3)_2$**	Manganese(II) chlorate
Sn^{4+}	**Br^-**	$SnBr_4$	**Tin(IV) bromide**
Co^{2+}	PO_4^{3-}	**$Co_3(PO_4)_2$**	**Cobalt(II) phosphate**
Hg_2^{2+}	I^-	**Hg_2I_2**	**Mercury(I) iodide**
Cu^+	**CO_3^{2-}**	Cu_2CO_3	**Copper(I) carbonate**
Li^+	**N^{3-}**	**Li_3N**	Lithium nitride
Al^{3+}	S^{2-}	**Al_2S_3**	**Aluminum sulfide**

2.118 a. $CO_2(s)$
b. NaCl
c. N_2O
d. $CaCO_3$
e. $NaHCO_3$
f. NH_3
g. H_2O
h. $Mg(OH)_2$
i. $MgSO_4{\cdot}7H_2O$

2.119 The change in energy is equal to the energy released. We call this ΔE. Similarly, Δm is the change in mass. Because $m = \dfrac{E}{c^2}$, we have

$$\Delta m = \frac{\Delta E}{c^2} = \frac{(1.715\times10^3\ \text{kJ})\left(\frac{1000\ \text{J}}{1\ \text{kJ}}\right)}{(2.998\times10^8\ \text{m/s})^2} = 1.908 \times 10^{-11}\ \text{kg} = \mathbf{1.908 \times 10^{-8}\ g}$$

Note that we need to convert kJ to J so that we end up with units of kg for the mass. $\left(1\ \text{J} = \dfrac{1\ \text{kg}\cdot\text{m}^2}{\text{s}^2}\right)$

We can add together the masses of hydrogen and oxygen to calculate the mass of water that should be formed.

$$12.096 \text{ g} + 96.000 \;=\; 108.096 \text{ g}$$

The predicted change (loss) in mass is only 1.908×10^{-8} g which is too small a quantity to measure. Therefore, for all practical purposes, the law of conservation of mass is assumed to hold for ordinary chemical processes.

2.120 a. Rutherford's experiment is described in detail in Section 2.2 of the text. From the average magnitude of scattering, Rutherford estimated the number of protons (based on electrostatic interactions) in the nucleus.

b. Assuming that the nucleus is spherical, the volume of the nucleus is:

$$V \;=\; \frac{4}{3}\pi r^3 \;=\; \frac{4}{3}\pi(3.04 \times 10^{-13}\text{ cm})^3 \;=\; 1.177 \times 10^{-37}\text{ cm}^3$$

The density of the nucleus can now be calculated.

$$\boldsymbol{d} \;=\; \frac{m}{V} \;=\; \frac{3.82 \times 10^{-23}\text{ g}}{1.177 \times 10^{-37}\text{ cm}^3} \;=\; \mathbf{3.25 \times 10^{14}\ g / cm^3}$$

To calculate the density of the space occupied by the electrons, we need both the mass of 11 electrons, and the volume occupied by these electrons.

The mass of 11 electrons is:

$$11 \text{ electrons} \times \frac{9.1094 \times 10^{-28}\text{ g}}{1 \text{ electron}} \;=\; 1.00203 \times 10^{-26}\text{ g}$$

The volume occupied by the electrons will be the difference between the volume of the atom and the volume of the nucleus. The volume of the nucleus was calculated above. The volume of the atom is calculated as follows:

$$186 \text{ pm} \times \frac{1 \times 10^{-12}\text{ m}}{1 \text{ pm}} \times \frac{1 \text{ cm}}{1 \times 10^{-2}\text{ m}} \;=\; 1.86 \times 10^{-8}\text{ cm}$$

$$V_{\text{atom}} \;=\; \frac{4}{3}\pi r^3 \;=\; \frac{4}{3}\pi(1.86 \times 10^{-8}\text{ cm})^3 \;=\; 2.695 \times 10^{-23}\text{ cm}^3$$

$$V_{\text{electrons}} \;=\; V_{\text{atom}} - V_{\text{nucleus}} \;=\; (2.695 \times 10^{-23}\text{ cm}^3) - (1.177 \times 10^{-37}\text{ cm}^3) \;=\; 2.695 \times 10^{-23}\text{ cm}^3$$

As you can see, the volume occupied by the nucleus is insignificant compared to the space occupied by the electrons.

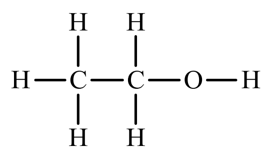
The density of the space occupied by the electrons can now be calculated.

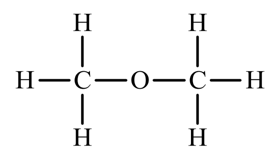
$$\boldsymbol{d} = \frac{m}{V} = \frac{1.00203 \times 10^{-26}\text{ g}}{2.695 \times 10^{-23}\text{ cm}^3} = \mathbf{3.72 \times 10^{-4}\ g/cm^3}$$

The above results do support Rutherford's model. Comparing the space occupied by the electrons to the volume of the nucleus, it is clear that most of the atom is empty space. Rutherford also proposed that the nucleus was a *dense* central core with most of the mass of the atom concentrated in it. Comparing the density of the nucleus with the density of the space occupied by the electrons also supports Rutherford's model.

2.121 The acids, from left to right, are chloric acid, nitrous acid, hydrocyanic acid, and sulfuric acid.

2.122 Two different structural formulas for the molecular formula C_2H_6O are:





The second hypothesis of Dalton's Atomic Theory states that compounds are composed of atoms of more than one element, and in any given compound, the same types of atoms are always present in the same relative numbers. Both of the above compounds are consistent with the second hypothesis.

2.123 a.

Ethane	Acetylene
2.65 g C	4.56 g C
0.665 g H	0.383 g H

Let's compare the ratio of the hydrogen masses in the two compounds. To do this, we need to start with the same mass of carbon. If we were to start with 4.56 g of C in ethane, how much hydrogen would combine with 4.56 g of carbon?

$$0.665\text{ g H} \times \frac{4.56\text{ g C}}{2.65\text{ g C}} = 1.14\text{ g H}$$

We can calculate the ratio of H in the two compounds.

$$\frac{1.14\text{ g}}{0.383\text{ g}} \approx 3$$

This is consistent with the Law of Multiple Proportions which states that if two elements combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers. In this case, the ratio of the masses of hydrogen in the two compounds is 3:1.

b. For a given amount of carbon, there is 3 times the amount of hydrogen in ethane compared to acetylene. Reasonable formulas would be:

Ethane	Acetylene
CH_3	CH
C_2H_6	C_2H_2

2.124 a. The following strategy can be used to convert from the volume of the Pt cube to the number of Pt atoms.

$$\text{cm}^3 \rightarrow \text{grams} \rightarrow \text{atoms}$$

$$1.0\ \text{cm}^3 \times \frac{21.45\ \text{g Pt}}{1\ \text{cm}^3} \times \frac{1\ \text{atom Pt}}{3.240 \times 10^{-22}\ \text{g Pt}} = \mathbf{6.6 \times 10^{22}\ Pt\ atoms}$$

b. Since 74 percent of the available space is taken up by Pt atoms, 6.6×10^{22} atoms occupy the following volume:

$$0.74 \times 1.0\ \text{cm}^3 = 0.74\ \text{cm}^3$$

We are trying to calculate the radius of a single Pt atom, so we need the volume occupied by a single Pt atom.

$$\text{volume Pt atom} = \frac{0.74\ \text{cm}^3}{6.6 \times 10^{22}\ \text{Pt atoms}} = 1.12 \times 10^{-23}\ \text{cm}^3/\text{Pt atom}$$

The volume of a sphere is $\frac{4}{3}\pi r^3$. Solving for the radius:

$$V = 1.12 \times 10^{-23}\ \text{cm}^3 = \frac{4}{3}\pi r^3$$

$$r^3 = 2.67 \times 10^{-24}\ \text{cm}^3$$

$$r = 1.4 \times 10^{-8}\ \text{cm}$$

Converting to picometers:

$$\textbf{radius Pt atom} = (1.4 \times 10^{-8}\ \text{cm}) \times \frac{0.01\ \text{m}}{1\ \text{cm}} \times \frac{1\ \text{pm}}{1 \times 10^{-12}\ \text{m}} = \mathbf{1.4 \times 10^2\ pm}$$

2.125 a. Assume that the nucleons (protons and neutrons) are hard objects of fixed size. Then the volume of the nucleus is well-approximated by the direct proportion $V = kA$, where A is the number of nucleons (mass number of the atom). For a spherical nucleus, then $V = kA = \frac{4}{3}\pi r^3$. Solving for *r*:

$$
\begin{aligned}
kA &= \frac{4}{3}\pi r^3 \\
\left(\frac{3}{4\pi}\right)kA &= r^3 \\
\left[\left(\frac{3}{4\pi}\right)kA\right]^{1/3} &= r \\
\left[\left(\frac{3}{4\pi}\right)k\right]^{1/3}\left(A^{1/3}\right) &= r \\
cA^{1/3} &= r \quad (c \text{ is a constant})
\end{aligned}
$$

b. For the volume calculation, use lithium-7 (*A* = 7).

$$V = \frac{4}{3}\pi r^3 = \frac{4}{3}\pi\left(r_0 A^{1/3}\right)^3 = \left(\frac{4}{3}\pi r_0^3\right)(A) = \left[\frac{4}{3}\pi\left(1.2\times10^{-15}\ \text{m}\right)^3\right](7) \approx 5.1\times10^{-44}\ \text{m}^3$$

c. Use $r = 152$ pm $= 152 \times 10^{-12}$ m for the atomic radius. Then, the atomic volume of lithium-7 is:

$$V = \frac{4}{3}\pi r^3 = \frac{4}{3}\pi\left(152\times10^{-12}\ \text{m}\right)^3 \approx 1.5\times10^{-29}\ \text{m}^3$$

The fraction of the atomic radius occupied by the nucleus is $\frac{5.1\times10^{-44}}{1.5\times10^{-29}} \approx 3.4\times10^{-15}$. This is consistent with Rutherford's discovery that the nucleus occupies a very small region within the atom.

2.126

Chapter 1

Chemistry: The Central Science

Practice Problems C

1.1 (iii)

1.2 (i) and (i)

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

1.4 physical: (iii), chemical: (i), (ii) is neither

1.5 12 blue cubes, infinite number of significant figures;

2×10^1 red spheres, one significant figure

1.6 2.4×10^2 lbs.

1.7 2.67 g/cm^3

1.8 (a) 4 red blocks/1 object

(b) 1 object/1 yellow block

(c) 2 white blocks/1 yellow block

(d) 1 yellow block/6 grey connectors

1.9 375 red bars; 3500 yellow balls

Questions and Problems

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

Matter is anything that has mass and occupies space.

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

1.3 **A hypothesis explains observations. A theory explains data from accumulated experiments and**

predicts related phenomena.

1.4 a. **Hypothesis** – This statement is an opinion.

b. **Law** – Newton's Law of Gravitation.

c. **Theory –** Atomic Theory.

1.5 a. **Law** – Newton's 2nd Law of Motion.

b. **Theory** – Big Bang Theory.

c. **Hypothesis** – It may be possible but we have no data to support this statement.

1.6 a. **O and H** b. **C and H** c. **H and Cl** d. **N**

1.7 a. **C and O** b. **F and H** c. **N and H** d. **O**

1.8 a. Matter is anything that has mass and occupies space. Examples include **air, seawater, concrete, an automobile, or a dog**.

b. A substance is a form of matter that has definite (constant) composition and distinct properties. Examples include **iron, silver, water, or sugar**.

c. A mixture is a combination of two or more substances in which the substances retain their distinct identities. Examples include **milk, salt water, air, or steel**.

1.9 Examples of homogeneous mixtures: **apple juice or root beer**.

Examples of heterogeneous mixtures: **chocolate chip cookie or vinaigrette salad dressing**.

1.10 Examples of elements (see the front cover for a complete list): **oxygen, platinum, sodium, cobalt.**

Examples of compounds: **sugar, salt, hemoglobin, citric acid.**

An element cannot be separated into simpler substances by chemical means. A compound can be separated into its constituent elements by a chemical reaction.

1.11 There are **118** known elements.

1.12 Li: **Lithium**

F: **Fluorine**

P: **Phosphorus**

Cu: **Copper**

As: **Arsenic**

Zn: **Zinc**

Cl: **Chlorine**

Pt: **Platinum**

Mg: **Magnesium**

U: **Uranium**

Al: **Aluminum**

Si: **Silicon**

Ne: **Neon**

1.13 a. **K** (potassium) d. **B** (boron) g. **S** (sulfur)

b. **Sn** (tin) e. **Ba** (barium) h. **A**r (argon)

c. **Cr** (chromium) f. **Pu** (plutonium) i. **Hg** (mercury)

1.14 a. hydrogen: **element** c. gold: **element**

b. water: **compound** d. sugar: **compound**

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

b. **element**

c. **compound**

d. **homogeneous mixture**

e. **heterogeneous mixture**

f. **homogeneous mixture**

g. **heterogeneous mixture**

1.16 a. **liquid** b. **gas** c. **mixture** d. **solid**

1.17 a. **element** b. **compound** c. **compound** d. **element**

1.18 a. Chemistry Units: **meter (m), centimeter (cm), millimeter (mm)**

SI Base Unit: **meter (m)**

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

SI Base Unit: **cubic meter (m^3)**

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

SI Base Unit: **kilogram (kg)**

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

SI Base Unit: **second (s)**

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

SI Base Unit: **kelvin (K)**

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

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

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

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

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

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

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

$$\text{weight on the moon } = (168 \text{ lbs on Earth})\left(\frac{1}{6}\right) = \mathbf{28\,lbs}$$

1.22 Kelvin is known as the absolute temperature scale, meaning the lowest possible temperature is 0 K.

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

$$\text{K} = °\text{C} + 273.15$$

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

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

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

1.23 **Strategy:** Use the density equation:

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

Solution:

$$\boldsymbol{d} = \frac{m}{V} = \frac{586\text{ g}}{188\text{ mL}} = \mathbf{3.12\ g\,/\,mL}$$

1.24

$$\textbf{mass of ethanol} = \frac{0.798\text{ g}}{1\text{ mL}} \times 17.4\text{ mL} = \mathbf{13.9\ g}$$

1.25 **Strategy:** Find the appropriate equations for converting between Fahrenheit and Celsius and between Celsius and Fahrenheit given in Section 1.3 of the text. Substitute the temperature values given in the problem into the appropriate equation.

Setup: Conversion from Fahrenheit to Celsius:

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

Conversion from Celsius to Fahrenheit:

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

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

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

c. $$°\text{C} = (102°\text{F} - 32°\text{F}) \times \frac{5°\text{C}}{9°\text{F}} = \mathbf{39°C}$$

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

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

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

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

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

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

Solution:

$$\text{volume of water} = V = \frac{m}{d} = \frac{2.50\ \text{g}}{0.992\ \text{g/mL}} = \mathbf{2.52\ mL}$$

1.28

$$\text{volume of platinum} = \frac{87.6\ \text{g}}{21.5\ \text{g/cm}^3} = \mathbf{4.07\ cm^3}$$

1.29 **Strategy:** Use the equation for converting °C to K. AMPS Solution Co

Setup: Conversion from Celsius to Kelvin:

$$\text{K} = °\text{C} + 273.15$$

Solution: a. **K** = 115.21°C + 273.15 = **388.36 K**

b. **K** = 37°C + 273 = $\mathbf{3.10 \times 10^2}$ **K**

c. **K** = 357°C + 273 = $\mathbf{6.30 \times 10^2}$ **K**

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

1.30 a. **°C** = K – 273 = 77 K – 273 = **–196°C** AMPS Solution Co

b. **°C** = 4.22 K – 273.15 = **–268.93°C**

c. **°C** = 600.61 K – 273.15 = **327.46°C**

1.31 **Qualitative data does not require explicit measurement. Quantitative data requires measurement and is expressed with a number.**

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

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

1.33 **An extensive property depends on the amount of substance present. An intensive property is independent of the amount of substance present.**

1.34 a. **extensive**

b. **extensive**

c. **intensive**

d. **extensive**

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

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

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

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

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

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

b. **Chemical property**. The fertilizer is consumed by the growing plants; it is turned into vegetable matter (different composition).

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

d. **Physical property**. The measurement of the densities of lead and aluminum does not change their composition.

e. **Chemical property**. When uranium undergoes nuclear decay, the products are chemically different substances.

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

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

c. **Physical Change.** The orange juice concentrate can be regenerated by evaporation of the water.

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

e. **Physical Change.** The salt can be recovered unchanged by evaporation.

1.38 Mass is extensive and additive: 44.3 + 115.2 = **159.5 g**

Temperature is intensive: **10°C**

Density is intensive: **1.00 g/mL**

1.39 Mass is extensive and additive: 37.2 + 62.7 = **99.9 g**

Temperature is intensive: **20°C**

Density is intensive: **11.35 g/cm^3**

1.40 a. **Exact.** The number of tickets is determined by counting.

b. **Inexact.** The volume must be measured.

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

d. **Inexact.** The mass of oxygen must be measured.

e. **Exact.** The number of days is a defined value.

1.41 **Using scientific notation avoids the ambiguity associated with trailing zeros.**

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

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

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

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

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

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

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

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

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

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

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

Solution: a. $1.52 \times 10^{-2} = \mathbf{0.0152}$

b. $7.78 \times 10^{-8} = \mathbf{0.0000000778}$

c. $1 \times 10^{-6} = \mathbf{0.000001}$

d. $1.6001 \times 10^{3} = \mathbf{1600.1}$

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

b. $\dfrac{79500}{2.5 \times 10^{2}} = \dfrac{7.95 \times 10^{4}}{2.5 \times 10^{2}} = \mathbf{3.2 \times 10^{2}}$

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

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

1.47 a. Addition using scientific notation.

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

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

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

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

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

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

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

b. Division using scientific notation.

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

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

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

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

$$6.53 \div 5.75 = 1.14$$

Subtract the exponents.

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

c. Subtraction using scientific notation.

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

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

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

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

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

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

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

d. Multiplication using scientific notation.

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

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

$$3.6 \times 3.6 = 13$$

Add the exponents.

$$13 \times 10^{-4 + (+6)} = 13 \times 10^2$$

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

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

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

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

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

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

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

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

1.50 a. **10.6 m** b. **0.79 g** c. **16.5 cm**

1.51 a. Division

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

Solution:

$$\frac{7.310\text{ km}}{5.70\text{ km}} = \mathbf{1.283}$$

The 3 (bolded) is a nonsignificant digit because the original number 5.70 only has three significant digits. Therefore, the answer has only three significant digits.

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

1.28

Think About It: Why are there no units?

b. Subtraction

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

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

$$\begin{array}{r} 0.00326\text{ mg} \\ -\ 0.0000788\text{ mg} \\ \hline \mathbf{0.0031812\ mg} \end{array}$$

The bold numbers are nonsignificant digits because the number 0.00326 has five digits to the right of the decimal point. Therefore, we carry five digits to the right of the decimal point in our answer.

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

$$\mathbf{0.00318\ mg = 3.18{\times}10^{-3}\ mg}$$

c. Addition

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

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

$$(0.402 \times 10^7\ \text{dm}) + (7.74 \times 10^7\ \text{dm}) = \mathbf{8.14 \times 10^7\ dm}$$

Since 7.74×10^7 has only two digits to the right of the decimal point, two digits are carried to the right of the decimal point in the final answer.

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

1.53 **Tailor Z's measurements are the most accurate. Tailor Y's measurements are the least accurate. Tailor X's measurements are the most precise. Tailor Y's measurements are the least precise.**

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

b. $$25.4\ \text{mg} \times \frac{1\times10^{-3}\ \text{g}}{1\ \text{mg}} \times \frac{1\times10^{-3}\ \text{kg}}{1\ \text{g}} = \mathbf{2.54\times10^{-5}}\ \text{kg}$$

c. $$556\ \text{mL} \times \frac{1\times10^{-3}\ \text{L}}{1\ \text{mL}} = \mathbf{0.556\ L}$$

d. $$\frac{10.6\ \text{kg}}{1\ \text{m}^3} \times \frac{1000\ \text{g}}{1\ \text{kg}} \times \left(\frac{1\times10^{-2}\ \text{m}}{1\ \text{cm}}\right)^3 = 0.0106\ \text{g}/\text{cm}^3$$

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

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

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

Solution:

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

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

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

Solution:

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

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

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

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

Solution:

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

Think About It: From the above conversion factors you can show that 1 m^3 = 1 × 10^3 L. Therefore, 7 m^3 would equal 7 × 10^3 L, which is close to the answer.

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

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

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

Solution:

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

Think About It: Does the answer seem reasonable? What number does the prefix μ represent? Should 28.3 μg be a very small mass?

1.56 $$\frac{1255\text{ m}}{1\text{ s}} \times \frac{1\text{ mi}}{1609\text{ m}} \times \frac{3600\text{ s}}{1\text{ h}} = \mathbf{2808\,mi / h}$$

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

Setup: The sequence of conversions is days → hours → minutes → seconds. Use the following conversion factors:

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

Solution:

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

Think About It: Does your answer seem reasonable? Should there be a very large number of seconds in 1 year?

1.58
$$(93 \times 10^6 \text{ mi}) \times \frac{1.609 \text{ km}}{1 \text{ mi}} \times \frac{1000 \text{ m}}{1 \text{ km}} \times \frac{1 \text{ s}}{3.00 \times 10^8 \text{ m}} \times \frac{1 \text{ min}}{60 \text{ s}} = \mathbf{8.3 \text{ min}}$$

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

Setup: Use the conversion factors:

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

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

Solution:

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

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

Setup: Use the conversion factor:

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

Solution:

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

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

Setup: Use the conversion factors:

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

Solution:

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

1.60

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

$$168\text{ lb} \times \frac{453.6\text{ g}}{1\text{ lb}} \times \frac{1\text{ kg}}{1000\text{ g}} = \mathbf{76.2\ kg}$$

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

Setup: Use the conversion factors:

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

Solution:

$$\frac{55\text{ mi}}{1\text{ h}} \times \frac{1609\text{ m}}{1\text{ mi}} \times \frac{1\text{ km}}{1000\text{ m}} = \mathbf{88\ km/h}$$

1.62 $\frac{62\text{ m}}{1\text{ s}} \times \frac{1\text{ mi}}{1609\text{ m}} \times \frac{3600\text{ s}}{1\text{ h}} = \mathbf{1.4 \times 10^2\ mph}$

1.63 **Strategy:** We seek to calculate the mass of Pb in a 6.0×10^3 g sample of blood. Lead is present in the blood at the rate of $0.62\text{ ppm} = \frac{0.62\text{ g Pb}}{1\times10^6\text{ g blood}}$. Use the rate to convert g blood → g Pb.

Setup: Be sure to set the conversion factor up so that g blood cancels.

Solution:

$$6.0 \times 10^3\text{ g of blood} \times \frac{0.62\text{ g Pb}}{1 \times 10^6\text{ g blood}} = \mathbf{3.7 \times 10^{-3}\ g\ Pb}$$

1.64 a. $$32.4\ \text{yd} \times \frac{36\ \text{in}}{1\ \text{yd}} \times \frac{2.54\ \text{cm}}{1\ \text{in}} = \mathbf{2.96 \times 10^{3}\ cm}$$

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

c. $$1.42\ \text{yr} \times \frac{365\ \text{day}}{1\ \text{yr}} \times \frac{24\ \text{h}}{1\ \text{day}} \times \frac{3600\ \text{s}}{1\ \text{h}} \times \frac{3.00 \times 10^{8}\ \text{m}}{1\ \text{s}} \times \frac{1\ \text{mi}}{1609\ \text{m}} = \mathbf{8.35 \times 10^{12}\ mi}$$

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

Setup: Use the conversion factor:

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

Solution:

$$185\ \text{nm} \times \frac{1 \times 10^{-9}\ \text{m}}{1\ \text{nm}} = \mathbf{1.85 \times 10^{-7}\ m}$$

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

Setup: Use the conversion factors:

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

Solution:

$$(4.5 \times 10^{9}\ \text{yr}) \times \frac{365\ \text{day}}{1\ \text{yr}} \times \frac{24\ \text{h}}{1\ \text{day}} \times \frac{3600\ \text{s}}{1\ \text{h}} = \mathbf{1.4 \times 10^{17}\ s}$$

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

Setup: Use the conversion factor:

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

Solution:

$$71.2\ \text{cm}^3 \times \left(\frac{0.01\ \text{m}}{1\ \text{cm}}\right)^3 = \mathbf{7.12 \times 10^{-5}\ m^3}$$

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

Setup: Use the conversion factors:

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

Solution:

$$88.6\ \text{m}^3 \times \left(\frac{1\ \text{cm}}{1 \times 10^{-2}\ \text{m}}\right)^3 \times \frac{1\ \text{L}}{1000\ \text{cm}^3} = \mathbf{8.86 \times 10^4\ L}$$

1.66

$$\textbf{density} = \frac{2.70\ \text{g}}{1\ \text{cm}^3} \times \frac{1\ \text{kg}}{1000\ \text{g}} \times \left(\frac{1\ \text{cm}}{0.01\ \text{m}}\right)^3 = \mathbf{2.70 \times 10^3\ kg / m^3}$$

1.67 **Strategy:** The given rate has units g/L and the desired units are g/cm^3. Use a conversion factor to convert $L \rightarrow cm^3$.

Setup: Use the conversion factor:

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

Solution:

$$\frac{0.625\ \text{g}}{1\ \text{L}} \times \frac{1\ \text{L}}{1000\ \text{cm}^3} = \mathbf{6.25 \times 10^{-4}\ g / cm^3}$$

1.68 a. Convert the dimensions of the room to dm:

$$17.6\ \text{m} \times \frac{10\ \text{dm}}{\text{m}} = 176\ \text{dm};\ 8.80\ \text{m} \times \frac{10\ \text{dm}}{\text{m}} = 88.0\ \text{dm};\ 2.64\ \text{m} \times \frac{10\ \text{dm}}{\text{m}} = 26.4\ \text{dm}$$

Multiply the room dimensions to get the volume of the room:

$$176\ \text{dm} \times 88.0\ \text{dm} \times 26.4\ \text{dm} = 4.089 \times 10^5\ \text{dm}^3 = 4.089 \times 10^5\ \text{L}$$

Because the concentration of CO in the room is 8.00×10^2 ppm, the volume that would be occupied under the same conditions by the CO alone is

$$4.089 \times 10^5\ \text{L} \times \frac{8.00 \times 10^2\ \text{L CO}}{1.00 \times 10^6\ \text{L total}} = \mathbf{327\ L}$$

b. $$\frac{0.050\ \text{mg}}{\text{m}^3} \times \frac{1\ \text{g}}{1 \times 10^3\ \text{mg}} \times \left(\frac{1\ \text{m}}{10\ \text{dm}}\right)^3 = 5.0 \times 10^{-8}\ \text{g/dm}^3 = \mathbf{5.0 \times 10^{-8}\ g/L}$$

c. $$\frac{120\ \text{mg}}{1\ \text{dL}} \times \frac{1 \times 10^3\ \mu\text{g}}{1\ \text{mg}} \times \frac{10\ \text{dL}}{1\ \text{L}} \times \frac{1\ \text{L}}{1000\ \text{mL}} = \mathbf{1.20 \times 10^3\ \mu g/mL}$$

1.69 **Strategy:** Use the equation $t = \dfrac{x^2}{2D}$ to convert the given distance ($x = 10\ \mu\text{m}$) to time t in seconds. Notice that both x and D contain distance units. But, for the given values, $x = 10\ \mu\text{m}$ and $D = 5.7 \times 10^{-7}\ \text{cm}^2/\text{s}$, the distance units are dissimilar and will not cancel. So, before calculating, convert the units of x from μm to cm.

Setup: Use the conversion factors:

$$\frac{1 \times 10^{-2}\ \text{m}}{1\ \text{cm}} \text{ and } \frac{1\ \mu\text{m}}{1 \times 10^{-6}\ \text{m}}$$

Solution:

$$10\ \mu\text{m} \times \frac{1 \times 10^{-6}\ \text{m}}{1\ \mu\text{m}} \times \frac{1\ \text{cm}}{1 \times 10^{-2}\ \text{m}} = 1 \times 10^{-3}\ \text{cm} = x \text{ in the equation.}$$

$$t = \left[\frac{(10^{-3}\ \text{cm})^2}{2 \times 5.7 \times 10^{-7}\ \text{cm}^2/\text{s}}\right] = \mathbf{0.88\ s}$$

1.70 Determine the volume occupied by the brain:

$$1\text{ kg}\times\frac{1000\text{ g}}{1\text{ kg}}\times\frac{1\text{ mL}}{1\text{ g}}=1000\text{ mL}$$

Next determine the volume occupied by a single cell:

$$\frac{1000\text{ mL}}{1\times10^{11}\text{ cells}}=1\times10^{-8}\text{ mL per cell}$$

Remember that 1 mL = 1 cm^3. The cube root of the volume per cell, expressed in cm^3, gives the length of a cubic cell side:

$$\sqrt[3]{1\times10^{-8}\text{ cm}^3}=2.2\times10^{-3}\text{cm}$$

Therefore, the length of the edge of each cube is **2×10^{-3} cm**.

The surface area of a single cell is $(2.2 \times 10^{-3}\text{ cm})^2 = 4.8 \times 10^{-6}\text{ cm}^2$.

The total surface area of 1×10^{11} cells in a single layer is

$$1\times10^{11}\text{ cells}\times\frac{4.8\times10^{-6}\text{cm}^2}{1\text{ cell}}=\mathbf{5\times10^5\text{ cm}^2}\times\left(\frac{1\text{ m}}{100\text{ cm}}\right)^2=\mathbf{50\text{ m}^2}$$

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

1.72 Volume of sample: V = 18.45 mL – 17.00 mL = 1.45mL

$$\text{Density: } d=\frac{m}{V}=\frac{13.2\text{ g}}{1.45\text{ mL}}=9.10\text{ g / mL}$$

1.73 a. **chemical** b. **chemical** c. **physical** d. **physical** e. **chemical**

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

$$\text{Volume }=(8.53\text{ cm})(2.4\text{ cm})(1.0\text{ cm})=2.0\times10^1\text{ cm}^3$$

$$d=\frac{m}{V}=\frac{52.7064\text{ g}}{2.0\times10^1\text{ cm}^3}=\mathbf{2.6\text{ g / cm}^3}$$

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

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

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

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

Solution:

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

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

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

$$V = s^3$$

Then, find the mass using the density equation:

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

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

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

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

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

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

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

Solution:

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

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

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

$$\text{volume of Hg filling cylinder} = \frac{105.5 \text{ g}}{13.6 \text{ g/cm}^3} = 7.757 \text{ cm}^3$$

Next, solve for the radius of the cylinder.

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

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

$$r = \sqrt{\frac{7.757 \text{ cm}^3}{\pi \times 12.7 \text{ cm}}} = 0.4409 \text{ cm}$$

The cylinder diameter equals $2r$.

$$\textbf{Cylinder diameter} = 2r = 2(0.4409 \text{ cm}) = \mathbf{0.882\ cm}$$

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

Setup: Solve the density equation for V:

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

Solution:

$$87.39 - 56.12 = 31.27 \text{ g water}$$

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

1.78 $\frac{343\text{ m}}{1\text{ s}} \times \frac{1\text{ mi}}{1609\text{ m}} \times \frac{3600\text{ s}}{1\text{ h}} = \mathbf{767\text{ mph}}$

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

Setup:

$$V = 260.5 - 242.0 = 18.50\text{ mL}$$

Solution:

$$d = \left(\frac{194.3\text{ g}}{18.50\text{ mL}}\right) = \mathbf{10.50\,g/mL}$$

The density of a solid is generally reported in g/cm^3. (1 mL = 1 cm^3) Therefore, the density is reported as **10.50 g/cm^3**.

Think About It: The volume of the water displaced must equal the volume of the piece of silver. If the silver did not sink, would you have been able to determine the volume of the piece of silver?

1.80 **The liquid must be *less dense* than the ice in order for the ice to sink. The temperature of the experiment must be maintained at or below 0°C to prevent the ice from melting.**

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

Setup:

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

Solution:

$$d = \frac{m}{V} = \frac{1.20 \times 10^4\text{ g}}{1.05 \times 10^3\text{ cm}^3} = \mathbf{11.4\text{ g/cm}^3}$$

1.82

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

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

1.83 **Strategy:** Use the conversion equation °F → °C or the conversion equation °C → °F. The solution below uses the conversion equation °C → °F.

Setup: Let *t* represent the common temperature. Substitute *t* into the conversion equation $°\text{F} = \left(°\text{C} \times \frac{9°\text{F}}{5°\text{C}}\right) + 32°\text{F}$ and solve for *t*.

Solution:

$$t = \frac{9}{5}t + 32°\text{F}$$

$$t - \frac{9}{5}t = 32°\text{F}$$

$$-\frac{4}{5}t = 32°\text{F}$$

$$\boldsymbol{t} \mathbf{= -40°F = -40°C}$$

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

$$\frac{195.6°\text{C}}{100°\text{S}}$$

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

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

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

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

1.85 **Strategy:** The volume of seawater is given. The strategy is to use the given conversion factors to convert L seawater → g seawater → g NaCl. This result can then be converted to kg NaCl and to tons

NaCl. Note that 3.1% NaCl by weight means 100 g seawater = 3.1 g NaCl.

Setup: Use the conversion factors:

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

Solution:

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

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

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

1.86 Volume = area × thickness.

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

$$\text{Volume}=\frac{\text{mass}}{\text{density}}=\frac{3.636\text{ g}}{2.699\text{ g / cm}^3}=1.3472\text{ cm}^3$$

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

$$1.000\text{ ft}^2\times\left(\frac{12\text{ in}}{1\text{ ft}}\right)^2\times\left(\frac{2.54\text{ cm}}{1\text{ in}}\right)^2=929.03\text{ cm}^2$$

$$\textbf{thickness}=\frac{\text{volume}}{\text{area}}=\frac{1.3472\text{ cm}^3}{929.03\text{ cm}^2}=1.450\times10^{-3}\text{ cm}=\mathbf{1.450\times10^{-2}\textbf{ mm}}$$

1.87 **Strategy:** Assume that the crucible is platinum. Calculate the volume of the crucible and then compare that to the volume of water that the crucible displaces.

Setup:

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

Solution:

$$\text{Volume of crucible} = \frac{860.2\text{ g}}{21.45\text{ g/cm}^3} = \mathbf{40.10\textbf{ cm}^3}$$

$$\text{Volume of water displaced} = \frac{(860.2 - 820.2)\text{g}}{0.9986\ \text{g/cm}^3} = \mathbf{40.1\ cm^3}$$

The volumes are the same (within experimental error), so **the density of the crucible is equal to the density of pure platinum**. Therefore, the crucible is probably made of platinum

1.88 Volume = surface area × depth

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

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

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

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

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

Setup: Conversion factor:

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

Solution:

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

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

Setup: Conversion factors:

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

Solution:

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

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

A troy ounce is heavier than an ounce.

1.90

$$\text{Volume of sphere} = \frac{4}{3}\pi r^3$$

$$\text{Volume} = \frac{4}{3}\pi\left(\frac{15\text{ cm}}{2}\right)^3 = 1.77 \times 10^3\text{ cm}^3$$

$$\text{mass} = \text{volume} \times \text{density} = (1.77 \times 10^3\text{ cm}^3) \times \frac{22.57\text{ g Os}}{1\text{ cm}^3} \times \frac{1\text{ kg}}{1000\text{ g}} = \mathbf{4.0 \times 10^1\ kg\ Os}$$

$$4.0 \times 10^1\text{ kg Os} \times \frac{2.205\text{ lb}}{1\text{ kg}} = 88\ \mathbf{lb\ Os}$$

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

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

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

Solution:

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

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

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

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

Solution:

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

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

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

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

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

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

Converting to nm:

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

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

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

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

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

Solution:

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

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

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

1.96 This problem is similar in concept to a limiting reactant problem. We need sets of coins with 3 quarters, 1 nickel, and 2 dimes. First, we need to find the total number of each type of coin.

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

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

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

Next, we need to find which coin limits the number of sets that can be assembled. For each set of coins, we need 2 dimes for every 1 nickel.

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

We do not have enough dimes.

For each set of coins, we need 2 dimes for every 3 quarters.

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

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

If we need 2 dimes per set, the number of sets that can be assembled is:

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

The mass of each set is:

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

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

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

1.97 **Strategy:** We are given a distance (1500 m) and a rate (3 minutes 43.13 seconds per mile). Use the rate as a conversion factor to convert m → mi → s.

Setup: Convert the time to run 1 mile to seconds:

$$3 \text{ min } 43.13 \text{ s} = 180 \text{ s} + 43.13 \text{ s} = 223.13 \text{ s}$$

Use the conversion factors: $\dfrac{1 \text{ mi}}{1609 \text{ m}}$ $\dfrac{223.13 \text{ s}}{1 \text{ mi}}$

Solution:

$$1500 \text{ m} \times \frac{1 \text{ mi}}{1609 \text{ m}} \times \frac{223.13 \text{ s}}{1 \text{ mi}} = \mathbf{208.0\ s} = \mathbf{3\ min\ 28.0\ s}$$

1.98

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

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

1.99 a. **homogeneous**

b. **heterogeneous**. The air will contain particulate matter, clouds, etc. This mixture is not homogeneous.

1.100

$$8 \times 10^4 \text{ tons Au} \times \frac{2000 \text{ lb Au}}{1 \text{ ton Au}} \times \frac{16 \text{ oz Au}}{1 \text{ lb Au}} \times \frac{28.35 \text{ g Au}}{1 \text{ oz Au}} \times \frac{\$1350}{31.103 \text{ g}} = \mathbf{\$3.2 \times 10^{12} \text{ or } \$3.2 \text{ trillion.}}$$

1.101 **Strategy:** Step 1: Use conversion factors to convert L seawater → mL seawater → g Au.

Step 2: Use conversion factors to convert g Au → dollars

Setup:

Step 1: Use the conversion factors: $\dfrac{1 \text{ mL seawater}}{0.001 \text{ L seawater}}$ $\dfrac{4.0 \times 10^{-12} \text{ g Au}}{1 \text{ mL seawater}}$

Step 2: Use the conversion factor: $\dfrac{\$1350}{31.103 \text{ g Au}}$

Solution:

$$(1.5 \times 10^{21} \text{ L seawater}) \times \frac{1 \text{ mL seawater}}{0.001 \text{ L seawater}} \times \frac{4.0 \times 10^{-12} \text{ g Au}}{1 \text{ mL seawater}} = \mathbf{6.0 \times 10^{12}\ g\ Au}$$

$$\textbf{value of gold} = 6.0 \times 10^{12} \text{ g Au} \times \frac{\$1350}{31.103 \text{ oz}} = \mathbf{\$2.6 \times 10^{14}}$$

Think About It: No one has become rich mining gold from the ocean, because the cost of recovering the gold would outweigh the price of the gold.

1.102 $$4.9\text{ g Fe} \times \frac{1.1 \times 10^{22}\text{ Fe atoms}}{1.0\text{ g Fe}} = \mathbf{5.4 \times 10^{22}\ Fe\ atoms}$$

1.103 **Strategy:** Use conversion factors to convert tons of earth → kg Si. Note that 0.50% crust by mass means 100 tons earth = 0.50 tons crust and that 27.2% Si by mass means 100 tons crust = 27.2 tons Si.

Setup: Conversion factors:

$$\frac{0.50\text{ ton crust}}{100\text{ ton earth}} \quad \frac{27.2\text{ ton Si}}{100\text{ ton crust}} \quad \frac{2000\text{ lb Si}}{1\text{ ton Si}} \quad \frac{453.6\text{ g Si}}{1\text{ lb Si}} \quad \frac{1\text{ kg Si}}{1000\text{ g Si}}$$

Solution:

$$5.9\times10^{21}\text{ ton earth}\times\frac{0.50\text{ ton crust}}{100\text{ ton earth}}\times\frac{27.2\text{ ton Si}}{100\text{ ton crust}}\times\frac{2000\text{ lb Si}}{1\text{ ton Si}}\times\frac{453.6\text{ g Si}}{1\text{ lb Si}}\times\frac{1\text{ kg Si}}{1000\text{ g Si}}$$

$$= 7.3\times10^{21}\text{ kg Si}$$

$$\textbf{mass of silicon in crust} = \mathbf{7.3 \times 10^{21}\ kg\ Si}$$

1.104 **Strategy:** The final cut results in two separate copper atoms. The length of a segment of wire consisting of a single copper atom is equal to the diameter (two times the radius) of a copper atom. We need to find the number of times the 0.1 m wire must be cut in half until the pieces that result from the final cut are each 2.6×10^{-10} m long.

Setup: Let *n* be the number of times we can cut the Cu wire in half. The original length is 10 cm or 0.1 m. We can write:

$$\left(\frac{1}{2}\right)^n \times 0.1\text{ m} = 2.6\times10^{-10}\text{ m}$$

$$\left(\frac{1}{2}\right)^n = 2.6\times10^{-9}\text{ m}$$

Solution: Taking the log of both sides of the equation:

$$n\log\left(\frac{1}{2}\right) = \log\left(2.6\times10^{-9}\ \text{m}\right)$$

$$\boldsymbol{n}\ \mathbf{=\ 29\ times}$$

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

Solution: From experiment one:

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

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

From the second experiment:

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

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

The volume of the ball bearing is obtained by difference.

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

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

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

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

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

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

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

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

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

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

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

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

1.107 **It would be more difficult to prove that the unknown substance is an element. Most compounds would decompose on heating, making them easy to identify. On heating, the compound HgO decomposes to elemental mercury (Hg) and oxygen gas (O_2).**

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

$$(2.0\times 10^4\text{ gallons H}_2\text{O})\times\frac{3.79\text{ L}}{1\text{ gallon}}\times\frac{1\text{ mL}}{0.001\text{ L}}\times\frac{1\text{ g}}{1\text{ mL}} = 7.58\times 10^7\text{ g H}_2\text{O}$$

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

$$(7.58\times 10^7\text{ g H}_2\text{O})\times\frac{1\text{ g chlorine}}{1\times 10^6\text{ g H}_2\text{O}} = 75.8\text{ g chlorine}$$

The chlorine solution is only 6 percent chlorine by mass. We can now calculate the volume of chlorine solution that must be added to the pool.

$$75.8\text{ g chlorine}\times\frac{100\%\text{ soln}}{6\%\text{ chlorine}}\times\frac{1\text{ mL soln}}{1\text{ g soln}} = \mathbf{1.3\times 10^3\ mL\ of\ chlorine\ solution}$$

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

Setup: Conversion factor:

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

Solution:

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

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

The volume of the cylinder is:

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

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

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

The density of the cylinder is:

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

Now, we can calculate the mass of the cylinder.

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

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

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

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

Setup: Use the conversion factors:

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

Solution:

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

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

First, convert from cubic feet to liters:

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

The mass of vanillin (in g) is:

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

The cost is:

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

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

Setup: Conversion factors:

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

Solution:

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

1.114 First, calculate the volume of 1 kg of seawater from the density and the mass. We chose 1 kg of seawater, because the problem gives the amount of Mg in every kg of seawater. The density of seawater is given in Problem 1.89.

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

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

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

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

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

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

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

1.115 **Strategy:** Use dimensional analysis. The conversions should convert the units "people" to the units "kg NaF". Since the number of conversion steps is large, divide the calculation into smaller steps.

Setup: Use the given conversion factors and also any others you may need from the inside back cover of the text.

Solution: The mass of water used by 50,000 people in 1 year is:

$$50,000\text{ people} \times \frac{150\text{ gal water}}{1\text{ person each day}} \times \frac{3.79\text{ L}}{1\text{ gal}} \times \frac{1000\text{ mL}}{1\text{ L}} \times \frac{1.0\text{ g H}_2\text{O}}{1\text{ mL H}_2\text{O}} \times \frac{365\text{ days}}{1\text{ yr}}$$
$$= 1.04 \times 10^{13}\text{ g H}_2\text{O/yr}$$

A concentration of 1 ppm of fluorine is needed. In other words, 1 g of fluorine is needed per million grams of water. NaF is 45.0% fluorine by mass. The amount of NaF needed per year in kg is:

$$(1.04 \times 10^{13}\text{ g H}_2\text{O}) \times \frac{1\text{ g F}}{10^6\text{ g H}_2\text{O}} \times \frac{100\%\text{ NaF}}{45\%\text{ F}} \times \frac{1\text{ kg}}{1000\text{ g}} = \mathbf{2.3 \times 10^4\ kg\ NaF}$$

An average person uses 150 gallons of water per day. This is equal to 569 L of water. If only 6 L of water is used for drinking and cooking, 563 L of water is used for purposes in which NaF is not necessary. Therefore the amount of NaF wasted is:

$$\frac{563\text{ L}}{569\text{ L}} \times 100\% = \mathbf{99\%}$$

1.116 $62\text{ kg} = 6.2 \times 10^4\text{ g}$

O: $(6.2 \times 10^4\text{ g})(0.65) = \mathbf{4.0 \times 10^4\ g\ O}$

C: $(6.2 \times 10^4\text{ g})(0.18) = \mathbf{1.1 \times 10^4\ g\ C}$

H: $(6.2 \times 10^4\text{ g})(0.10) = \mathbf{6.2 \times 10^3\ g\ H}$

N: $(6.2 \times 10^4\text{ g})(0.03) = \mathbf{2 \times 10^3\ g\ N}$

Ca: $(6.2 \times 10^4\text{ g})(0.016) = \mathbf{9.9 \times 10^2\ g\ Ca}$

P: $(6.2 \times 10^4\text{ g})(0.012) = \mathbf{7.4 \times 10^2\ g\ P}$

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

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

Solution:

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

Since there are 12 breaths per min,

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

1.118 a.

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

b.

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

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

Setup:

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

For the Fahrenheit thermometer, we expect:

$$|\text{true value} - \text{experimental value}| = 0.056°\text{C}$$

For the Celsius thermometer, we expect:

$$|\text{true value} - \text{experimental value}| = 0.1°\text{C}$$

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

Solution:

For the Fahrenheit thermometer, **percent error** $= \dfrac{0.056°\text{C}}{38.9°\text{C}} \times 100\% = \mathbf{0.1\%}$

For the Celsius thermometer, **percent error** $= \dfrac{0.1°\text{C}}{38.9°\text{C}} \times 100\% = \mathbf{0.3\%}$

Think About It: Which thermometer is more precise?

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

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

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

We can now calculate the mass of carbon dioxide by difference.

$$\text{Mass of CO}_2\text{ released} = 46.928\text{ g} - 46.699\text{ g} = 0.229\text{ g}$$

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

$$\text{Volume of CO}_2\text{ released} = 0.229\text{ g} \times \frac{1\text{ L}}{1.81\text{ g}} = \mathbf{0.127\text{ L}}$$

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

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

Converting the radius and height to cm gives:

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

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

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

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

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

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