Appendices

Practice!

Chapter 2

Section 2-1

1. The density of a substance is 4.8 g/mL. What is the volume of a sample that is 19.2 g?

2. A 2.00-mL sample of substance A has a density of 18.4 g/mL and a 5.00-mL sample of substance B has a density of 35.5 g/mL. Do you have an equal mass of substances A and B?

Section 2-2

3. Express the following quantities in scientific notation.

a. 5 453 000 m
b. 300.8 kg
c. 0.005 36 ng
d. 0.012 032 5 km
e. 34 800 s
f. 332 080 000 cm
g. 0.000 238 3 ms
h. 0.3048 mL

4. Solve the following problems. Express your answers in scientific notation.

a. 3×10^{2} m + 5×10^{2} m
b. 8×10^{-5} m + 4×10^{-5} m
c. 6.0×10^{5} m + 2.38×10^{6} m
d. 2.3×10^{-3} L + 5.78×10^{-2} L
e. 2.56×10^{2} g − 1.48×10^{2} g
f. 5.34×10^{-3} L − 3.98×10^{-3} L
g. 7.623×10^{5} nm − 8.32×10^{4} nm
h. 9.052×10^{-2} s − 3.61×10^{-3} s

5. Solve the following problems. Express your answers in scientific notation.

a. $(8 \times 10^{3}$ m$) \times (1 \times 10^{5}$ m$)$
b. $(4 \times 10^{2}$ m$) \times (2 \times 10^{4}$ m$)$
c. $(5 \times 10^{-3}$ m$) \times (3 \times 10^{4}$ m$)$
d. $(3 \times 10^{-4}$ m$) \times (3 \times 10^{-2}$ m$)$
e. $(8 \times 10^{4}$ g$) \div (4 \times 10^{3}$ mL$)$
f. $(6 \times 10^{-3}$ g$) \div (2 \times 10^{-1}$ mL$)$
g. $(1.8 \times 10^{-2}$ g$) \div (9 \times 10^{-5}$ mL$)$
h. $(4 \times 10^{-4}$ g$) \div (1 \times 10^{3}$ mL$)$

6. Convert the following as indicated.

a. 96 kg to g
b. 155 mg to g
c. 15 cg to kg
d. 584 μs to s
e. 188 dL to L
f. 3600 m to km
g. 24 g to pg
h. 85 cm to nm

7. How many minutes are there in 5 days?

8. A car is traveling at 118 km/h. What is its speed in Mm/h?

Section 2-3

9. Three measurements of 34.5 m, 38.4 m, and 35.3 m are taken. If the accepted value of the measurement is 36.7 m, what is the percent error for each measurement?

10. Three measurements of 12.3 mL, 12.5 mL, and 13.1 mL are taken. The accepted value for each measurement is 12.8 mL. Calculate the percent error for each measurement.

11. Determine the number of significant figures in each measurement.
a. 340 438 g
b. 87 000 ms
c. 4080 kg
d. 961 083 110 m
e. 1.040 s
f. 0.0483 m
g. 0.2080 mL
h. 0.000 048 1 g

12. Write the following in three significant figures.
a. 0.003 085 0 km
b. 3.0823 g
c. 5808 mL
d. 34.654 mg

13. Write the answers in scientific notation.
a. 0.005 832g
b. 386 808 ns
c. 0.000 580 0 km
d. 2086 L

14. Use rounding rules when you complete the following.
a. 34.3 m + 35.8 m + 33.7 m
b. 0.056 kg + 0.0783 kg + 0.0323 kg
c. 309.1 mL + 158.02 mL + 238.1 mL
d. 1.03 mg + 2.58 mg + 4.385 mg
e. 8.376 km − 6.153 km
f. 34.24 s −12.4 s
g. 804.9 dm − 342.0 dm
h. 6.38×10^2 m − 1.57×10^2 m

15. Complete the following calculations. Round off the answers to the correct number of significant figures.
a. 34.3 cm × 12 cm
b. 0.054 mm × 0.3804 mm
c. 45.1 km × 13.4 km
d. 45.5g ÷ 15.5 mL
e. 35.43 g ÷ 24.84 mL
f. 0.0482 g ÷ 0.003 146 mL

Chapter 3

Section 3-2

1. A 3.5-kg iron shovel is left outside through the winter. The shovel, now orange with rust, is rediscovered in the spring. Its mass is 3.7 kg. How much oxygen combined with the iron?

2. When 5.0 g of tin reacts with hydrochloric acid, the mass of the products, tin chloride and hydrogen, totals 8.1 g. How many grams of hydrochloric acid were used?

Section 3-3

3. A compound is analyzed and found to be 50.0% sulfur and 50.0% oxygen. If the total amount of the sulfur oxide compound is 12.5 g, how many grams of sulfur are there?

4. Two unknown compounds are analyzed. Compound I contain 5.63 g of tin and 3.37 g of chlorine, while compound II contains 2.5 g of tin and 2.98 g of chlorine. Are the compounds the same?

Chapter 4

Section 4-3

1. How many protons and electrons are in each of the following atoms?
a. gallium
b. silicon
c. cesium
d. calcium
e. molybdenum
f. titanium

2. What is the atomic number of each of the following elements?
a. an atom that contains 37 electrons
b. an atom that contains 72 protons

c. an atom that contains 1 electron
d. an atom that contains 85 protons

3. Use the periodic table to write the name and the symbol for each element identified in question 2.

4. An isotope of copper contains 29 electrons, 29 protons, and 36 neutrons. What is the mass number of this isotope?

5. An isotope of uranium contains 92 electrons and 144 neutrons. What is the mass number of this isotope?

6. Use the periodic table to write the symbols for each of the following elements. Then, determine the number of electrons, protons, and neutrons each contains.
   a. yttrium-88
   b. arsenic-75
   c. xenon-129
   d. bromine-79
   e. gold-197
   f. helium-4

7. An element has two naturally occurring isotopes: ^{14}X and ^{15}X. ^{14}X has a mass of 14.003 07 amu and a relative abundance of 99.63%. ^{15}X has a mass of 15.000 11 amu and a relative abundance of 0.37%. Identify the unknown element.

8. Silver has two naturally occurring isotopes. Ag-107 has an abundance of 51.82% and a mass of 106.9 amu. Ag-109 has a relative abundance of 48.18% and a mass of 108.9 amu. Calculate the atomic mass of silver.

Chapter 5

Section 5-1

1. What is the frequency of an electromagnetic wave that has a wavelength of 4.55×10^{-3} m? 1.00×10^{-12} m?

2. Calculate the wavelength of an electromagnetic wave with a frequency of 8.68×10^{16} Hz; 5.0×10^{14} Hz; 1.00×10^{6} Hz.

3. What is the energy of a quantum of visible light having a frequency of 5.45×10^{14} s^{-1}?

4. An X ray has a frequency of 1.28×10^{18} s^{-1}. What is the energy of a quantum of the X ray?

Section 5-3

5. Write the ground-state electron configuration for the following.
   a. nickel
   b. cesium
   c. boron
   d. krypton

6. What element has the following ground-state electron configuration $[He]2s^2$? $[Xe]6s^2 4f^{14} 5d^{10} 6p^1$?

7. Which element in period 4 has four electrons in its electron-dot structure?

8. Which element in period 2 has six electrons in its electron-dot structure?

9. Draw the electron-dot structure for each element in question 5.

Chapter 6

Section 6-2

1. Identify the group, period, and block of an atom with the following electron configuration.
a. $[He]2s^22p^1$ **b.** $[Kr]5s^24d^5$ **c.** $[Xe]6s^25f^{14}6d^5$

2. Write the electron configuration for the element fitting each of the following descriptions.
a. The noble gas in the first period.
b. The group 4B element in the fifth period.
c. The group 4A element in the sixth period.
d. The group 1A element in the seventh period.

Section 6-3

3. Using the periodic table and not **Figure 6-11,** rank each main group element in order of increasing size.
a. calcium, magnesium, and strontium
b. oxygen, lithium, and fluorine
c. fluorine, cesium, and calcium
d. selenium, chlorine, tellurium
e. iodine, krypton, and beryllium

Chapter 8

Section 8-2

1. Explain the formation of an ionic compound from zinc and chlorine.

2. Explain the formation of an ionic compound from barium and nitrogen.

Section 8-3

3. Write the chemical formula of an ionic compound composed of the following ions.
a. calcium and arsenide
b. iron(III) and chloride
c. magnesium and sulfide
d. barium and iodide
e. gallium and phosphide

4. Determine the formula for ionic compounds composed of the following ions.
a. copper(II) and acetate
b. ammonium and phosphate
c. calcium and hydroxide
d. gold(III) and cyanide

5. Name the following compounds.
a. $Co(OH)_2$
b. $Ca(ClO_3)_2$
c. Na_3PO_4
d. $K_2Cr_2O_7$
e. SrI_2
f. HgF_2

Chapter 9

Section 9-1

1. Draw the Lewis structure for the following molecules.
a. CCl_2H_2
b. HF
c. PCl_3
d. CH_4

Section 9-2

2. Name the following binary compounds.
a. S_4N_2
b. OCl_2
c. SF_6
d. NO
e. SiO_2
f. IF_7

3. Name the following acids: H_3PO_4, HBr, HNO_3.

Section 9-3

4. Draw the Lewis structure for each of the following.
   - **a.** CO
   - **b.** CH_2O
   - **c.** N_2O
   - **d.** OCl_2
   - **e.** SiO_2
   - **f.** $AlBr_3$
5. Draw the Lewis resonance structure for CO_3^{2-}.
6. Draw the Lewis resonance structure for $CH_3CO_2^-$.
7. Draw the Lewis structure for NO and IF_4^-.

Section 9-4

8. Determine the molecular geometry, bond angles, and hybrid of each molecule in question 4.

Section 9-5

9. Determine whether each of the following molecules is polar or nonpolar.
   - **a.** CH_2O
   - **b.** BF_3
   - **c.** SiH_4
   - **d.** H_2S

Chapter 10

Section 10-1 **Write skeleton equations for the following reactions.**

1. Solid barium and oxygen gas react to produce solid barium oxide.
2. Solid iron and aqueous hydrogen sulfate react to produce aqueous iron(III) sulfate and gaseous hydrogen.

Write balanced chemical equations for the following reactions.

3. Liquid bromine reacts with solid phosphorus (P_4) to produce solid diphosphorus pentabromide.
4. Aqueous lead(II) nitrate reacts with aqueous potassium iodide to produce solid lead(II) iodide and aqueous potassium nitrate.
5. Solid carbon reacts with gaseous fluorine to produce gaseous carbon tetrafluoride.
6. Aqueous carbonic acid reacts to produce liquid water and gaseous carbon dioxide.
7. Gaseous hydrogen chloride reacts with gaseous ammonia to produce solid ammonium chloride.
8. Solid copper(II) sulfide reacts with aqueous nitric acid to produce aqueous copper(II) sulfate, liquid water, and nitrogen dioxide gas.

Section 10-2 **Classify each of the following reactions in as many classes as possible.**

9. $2Mo(s) + 3O_2(g) \rightarrow 2MoO_3(s)$
10. $N_2H_4(l) + 3O_2(g) \rightarrow 2NO_2(g) + 2H_2O(l)$

Write balanced chemical equations for the following decomposition reactions.

11. Aqueous hydrogen chlorite decomposes to produce water and gaseous chlorine(III) oxide.
12. Calcium carbonate(s) decomposes to produce calcium oxide(s) and carbon dioxide(g).

Use the activity series to predict whether each of the following single-replacement reactions will occur:

13. $Al(s) + FeCl_3(aq) \rightarrow AlCl_3(aq) + Fe(s)$
14. $Br_2(l) + 2LiI(aq) \rightarrow 2LiBr(aq) + I_2(aq)$
15. $Cu(s) + MgSO_4(aq) \rightarrow Mg(s) + CuSO_4(aq)$

Write chemical equations for the following chemical reactions:

16. Bismuth(III) nitrate(aq) reacts with sodium sulfide(aq) yielding bismuth(III) sulfide(s) plus sodium nitrate(aq).

17. Magnesium chloride(aq) reacts with potassium carbonate(aq) yielding magnesium carbonate(s) plus potassium chloride(aq).

Section 10-3 **Write net ionic equations for the following reactions.**

18. Aqueous solutions of barium chloride and sodium fluoride are mixed to form a precipitate of barium fluoride.

19. Aqueous solutions of copper(I) nitrate and potassium sulfide are mixed to form insoluble copper(I) sulfide.

20. Hydrobromic acid reacts with aqueous lithium hydroxide

21. Perchloric acid reacts with aqueous rubidium hydroxide

22. Nitric acid reacts with aqueous sodium carbonate.

23. Hydrochloric acid reacts with aqueous lithium cyanide.

Chapter 11

Section 11-1

1. Determine the number of atoms in 3.75 mol Fe.

2. Calculate the number of formula units in 12.5 mol $CaCO_3$.

3. How many moles of $CaCl_2$ contains 1.26×10^{24} formula units $CaCl_2$?

4. How many moles of Ag contains 4.59×10^{25} atoms Ag?

Section 11-2

5. Determine the mass in grams of 0.0458 moles of sulfur.

6. Calculate the mass in grams of 2.56×10^{-3} moles of iron.

7. Determine the mass in grams of 125 mol of neon.

8. How many moles of titanium are contained in 71.4 g?

9. How many moles of lead are equivalent to 9.51×10^{3} g Pb?

10. Determine the number of moles of arsenic in 1.90 g As.

11. Determine the number of atoms in 4.56×10^{-2} g of sodium.

12. How many atoms of gallium are in 2.85×10^{3} g of gallium?

13. Determine the mass in grams of 5.65×10^{24} atoms Se.

14. What is the mass in grams of 3.75×10^{21} atoms Li?

Section 11-3

15. How many moles of each element is in 0.0250 mol K_2CrO_4.

16. How many moles of ammonium ions are in 4.50 mol $(NH_4)_2CO_3$?

17. Determine the molar mass of silver nitrate.

18. Calculate the molar mass of acetic acid (CH_3COOH).

19. Determine the mass of 8.57 mol of sodium dichromate ($Na_2Cr_2O_7$).

20. Calculate the mass of 42.5 mol of potassium cyanide.

21. Determine the number of moles present in 456 g $Cu(NO_3)_2$.

22. Calculate the number of moles in 5.67 g potassium hydroxide.

23. Calculate the number of each atom in 40.0 g of methanol (CH_3OH).

Practice!

24. What mass of sodium hydroxide contains 4.58×10^{23} formula units?

Section 11-4 25. What is the percent by mass of each element in sucrose ($C_{12}H_{22}O_{11}$)?

26. Which of the following compounds has a greater percent by mass of chromium, K_2CrO_4 or $K_2Cr_2O_7$?

27. Analysis of a compound indicates the percent composition 42.07% Na, 18.89% P, and 39.04% O. Determine its empirical formula.

28. A colorless liquid was found to contain 39.12% C, 8.76% H, and 52.12% O. Determine the empirical formula of the substance.

29. Analysis of a compound used in cosmetics reveals the compound contains 26.76% C, 2.21% H, 71.17% O and has a molar mass of 90.04 g/mol. Determine the molecular formula for this substance.

30. Eucalyptus leaves are the food source for panda bears. Eucalyptol is an oil found in these leaves. Analysis of eucalyptol indicates it has a molar mass of 154 g/mol and contains 77.87% C, 11.76% H, and 10.37% O. Determine the molecular formula of eucalyptol.

31. Beryl is a hard mineral which occurs in a variety of colors. A 50.0-g sample of beryl contains 2.52 g Be, 5.01 g Al, 15.68 g Si, and 26.79g O. Determine its empirical formula.

32. Analysis of a 15.0-g sample of a compound used to leach gold from low grade ores is 7.03 g Na, 3.68 g C, and 4.29 g N. Determine the empirical formula for this substance.

Section 11-5 33. Analysis of a hydrate of iron(III) chloride revealed that in a 10.00-g sample of the hydrate, 6.00 g is anhydrous iron(III) chloride and 4.00 g is water. Determine the formula and name of the hydrate.

34. When 25.00 g of a hydrate of nickel(II) chloride was heated, 11.37 g of water were released. Determine the name and formula of the hydrate.

Chapter 12

Section 12-1 **Interpret the following balanced chemical equation in terms of particles, moles, and mass.**

1. $Mg + 2HCl \rightarrow MgCl_2 + H_2$
2. $2Al + 3CuSO_4 \rightarrow Al_2(SO_4)_3 + 3Cu$
3. Write and balance the equation for the decomposition of aluminum carbonate. Determine the possible mole ratios.
4. Write and balance the equation for the formation of magnesium hydroxide and hydrogen from magnesium and water. Determine the possible mole ratios.

Section 12-2

5. Some antacid tablets contain aluminum hydroxide. The aluminum hydroxide reacts with stomach acid according to the equation: $Al(OH)_3 + 3HCl \rightarrow AlCl_3 + 3H_2O$. Determine the moles of acid neutralized if a tablet contains 0.200 mol $Al(OH)_3$.
6. Chromium reacts with oxygen according to the equation: $4Cr + 3O_2 \rightarrow 2Cr_2O_3$. Determine the moles of chromium(III) oxide produced when 4.58 moles of chromium is allowed to react.

7. Space vehicles use solid lithium hydroxide to remove exhaled carbon dioxide according to the equation: $2LiOH + CO_2 \rightarrow Li_2CO_3 + H_2O$. Determine the mass of carbon dioxide removed if the space vehicle carries 42.0 mol LiOH.

8. Some of the sulfur dioxide released into the atmosphere is converted to sulfuric acid according to the equation $2SO_2 + 2H_2O + O_2 \rightarrow 2H_2SO_4$. Determine the mass of sulfuric acid formed from 3.20 moles of sulfur dioxide.

9. How many grams of carbon dioxide are produced when 2.50 g of sodium hydrogen carbonate react with excess citric acid according to the equation: $3NaHCO_3 + H_3C_6H_5O_7 \rightarrow Na_3C_6H_5O_7 + 3CO_2 + 3H_2O$.

10. Aspirin ($C_9H_8O_4$) is produced when salicylic acid ($C_7H_6O_3$) reacts with acetic anhydride ($C_4H_6O_3$) according to the equation: $C_7H_6O_3 + C_4H_6O_3 \rightarrow C_9H_8O_4 + HC_2H_3O_2$. Determine the mass of aspirin produced when 150.0 g of salicylic acid reacts with an excess of acetic anhydride.

Section 12-3

11. Chlorine reacts with benzene to produce chlorobenzene and hydrogen chloride, $Cl_2 + C_6H_6 \rightarrow C_6H_5Cl + HCl$. Determine the limiting reactant if 45.0 g of benzene reacts with 45.0 g of chlorine, the mass of the excess reactant after the reaction is complete, and the mass of chlorobenzene produced.

12. Nickel reacts with hydrochloric acid to produce nickel(II) chloride and hydrogen according to the equation $Ni + 2HCl \rightarrow NiCl_2 + H_2$. If 5.00 g Ni and 2.50 g HCl react, determine the limiting reactant, the mass of the excess reactant after the reaction is complete, and the mass of nickel(II) chloride produced.

Section 12-4

13. Tin(IV) iodide is prepared by reacting tin with iodine. Write the balanced chemical equation for the reaction. Determine the theoretical yield if a 5.00-g sample of tin reacts in an excess of iodine. Determine the percent yield, if 25.0 g SnI_4 was actually recovered.

14. Gold is extracted from gold bearing rock by adding sodium cyanide in the presence of oxygen and water, according to the reaction, $4\ Au\ (s) + 8\ NaCN\ (aq) + O_2\ (g) + 2H_2O(l) \rightarrow 4\ NaAu(CN)_2\ (aq) + NaOH\ (aq)$. Determine the theoretical yield of $NaAu(CN)_2$ if 1000.0 g of gold bearing rock is used which contains 3.00% gold by mass. Determine the percent yield of $NaAu(CN)_2$ if 38.790 g $NaAu(CN)_2$ is recovered.

Chapter 13

Section 13-1

1. Calculate the ratio of effusion rates for methane (CH_4) and nitrogen.

2. Calculate the molar mass of butane. Butane's rate of diffusion is 3.8 times slower than that of helium.

Section 13-2

3. What is the total pressure in a canister that contains oxygen gas at a partial pressure of 804 mm Hg, nitrogen at a partial pressure of 220 mm Hg, and hydrogen at a partial pressure of 445 mm Hg?

4. Calculate the partial pressure of neon in a flask that has a total pressure of 1.87 atm. The flask contains krypton at a partial pressure of 0.77 atm and helium at a partial pressure of 0.62 atm.

Chapter 14

Section 14-1

1. The pressure of air in a 2.25-L container is 1.20 atm. What is the new pressure if the sample is transferred to a 6.50-L container? Temperature is constant.
2. The volume of a sample of hydrogen gas at 0.997 atm is 5.00 L. What will be the new volume if the pressure is decreased to 0.977 atm? Temperature is constant.
3. A gas at 55.0°C occupies a volume of 3.60 L. What volume will it occupy at 30.0°C? Pressure is constant.
4. The volume of a gas is 0.668 L at 66.8°C. At what Celsius temperature will the gas have a volume of 0.942 L, assuming the pressure remains constant?
5. The pressure in a bicycle tire is 1.34 atm at 33.0°C. At what temperature will the pressure inside the tire be 1.60 atm? Volume is constant.
6. If a sample of oxygen gas has a pressure of 810 torr at 298 K, what will be its pressure if its temperature is raised to 330 K?
7. Air in a tightly sealed bottle has a pressure of 0.978 atm at 25.5°C. What will its pressure be if the temperature is raised to 46.0°C?

Section 14-2

8. Hydrogen gas at a temperature of 22.0°C that is confined in a 5.00-L cylinder exerts a pressure of 4.20 atm. If the gas is released into a 10.0-L reaction vessel at a temperature of 33.6°C, what will be the pressure inside the reaction vessel?
9. A sample of neon gas at a pressure of 1.08 atm fills a flask with a volume of 250 mL at a temperature of 24.0°C. If the gas is transferred to another flask at 37.2°C at a pressure of 2.25 atm, what is the volume of the new flask?
10. What volume of beaker contains exactly 2.23×10^{-2} mol of nitrogen gas at STP?
11. How many moles of air are in a 6.06-L tire at STP?
12. How many moles of oxygen are in a 5.5-L canister at STP?
13. What mass of helium is in a 2.00-L balloon at STP?
14. What volume will 2.3 kg of nitrogen gas occupy at STP?

Section 14-3

15. Calculate the number of moles of gas that occupy a 3.45-L container at a pressure of 150 kPa and a temperature of 45.6°C.
16. What is the pressure in torr that a 0.44-g sample of carbon dioxide gas will exert at a temperature of 46.2°C when it occupies a volume of 5.00 L?
17. What is the molar mass of a gas that has a density of 1.02 g/L at 0.990 atm pressure and 37°C?

18. Calculate the grams of oxygen gas present in a 2.50-L sample kept at 1.66 atm pressure and a temperature of 10.0°C.

19. What volume of oxygen gas is needed to completely combust 0.202 L of butane (C_4H_{10}) gas?

20. Determine the volume of methane (CH_4) gas needed to react completely with 0.660 L of O_2 gas to form methanol (CH_3OH).

Section 14-4

21. Calculate the mass of hydrogen peroxide needed to obtain 0.460 L of oxygen gas at STP. $2H_2O_2(aq) \longrightarrow 2H_2O(l) + O_2(g)$

22. When potassium chlorate is heated in the presence of a catalyst such as manganese dioxide, it decomposes to form solid potassium chloride and oxygen gas: $2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$. How many liters of oxygen will be produced at STP if 1.25 kg of potassium chlorate decomposes completely?

Chapter 15

Section 15-1

1. Calculate the mass of gas dissolved at 150.0 kPa, if 0.35 g of the gas dissolves in 2.0 L of water at 30.0 kPa.

2. At which depth, 33 ft. or 133 ft, will a scuba diver have more nitrogen dissolved in the bloodstream?

Section 15-2

3. What is the percent by mass of a sample of ocean water that is found to contain 1.36 grams of magnesium ions per 1000 g?

4. What is the percent by mass of iced tea containing 0.75 g of aspartame in 250 g of water?

5. A bottle of hydrogen peroxide is labeled 3%. If you pour out 50 mL of hydrogen peroxide solution, what volume is actually hydrogen peroxide?

6. If 50 mL of pure acetone is mixed with 450 mL of water, what is the percent volume?

7. Calculate the molarity of 1270 g K_3PO_4 in 4.0 L aqueous solution.

8. What is the molarity of 90.0 g NH_4Cl in 2.25 L aqueous solution?

9. Which is more concentrated, 25 g NaCl dissolved in 500 mL of water or a 10% solution of NaCl (percent by mass)?

10. Calculate the mass of NaOH required to prepare a 0.343*M* solution dissolved in 2500 mL of water?

11. Calculate the volume required to dissolve 11.2 g $CuSO_4$ to prepare a 0.140*M* solution.

12. How would you prepare 500 mL of a solution that has a new concentration of 4.5*M* if the stock solution is 11.6*M*?

13. Caustic soda is 19.1*M* NaOH and is diluted for household use. What is the household concentration if 10 mL of the concentrated solution is diluted to 400 mL?

14. What is the molality of a solution containing 63.0 g HNO_3 in 0.500 kg of water?

15. What is the molality of an acetic acid solution containing 0.500 mole of $HC_2H_3O_2$ in 0.800 kg of water?

16. What mass of ethanol (C_2H_5OH) will be required to prepare a 2.00*m* solution in 8.00 kg of water?

17. Determine the mole fraction of nitrogen in a gas mixture containing 0.215 mol N_2, 0.345 mol O_2, 0.023 mol CO_2, and 0.014 mol SO_2. What is the mole fraction of N_2?

18. A necklace contains 4.85 g of gold, 1.25 g of silver, and 2.40 g of copper. What is the mole fraction of each metal?

Section 15-3

19. Calculate the freezing point and boiling point of a solution containing 6.42 g of sucrose ($C_{12}H_{22}O_{11}$) in 100.0 g of water.

20. Calculate the freezing point and boiling point of a solution containing 23.7 g copper(II) sulfate in 250.0 g of water.

21. Calculate the freezing point and boiling point of a solution containing 0.15 mol of the molecular compound naphthalene in 175 g of benzene (C_6H_6).

Chapter 16

Section 16-1

1. What is the equivalent in joules of 126 Calories?

2. Convert 455 kilojoules to kilocalories.

3. How much heat is required to warm 122 g of water by 23.0°C?

4. The temperature of 55.6 grams of a material decreases by 14.8°C when it loses 3080 J of heat. What is its specific heat?

5. What is the specific heat of a metal if the temperature of a 12.5-g sample increases from 19.5°C to 33.6°C when it absorbs 37.7 J of heat?

Section 16-2

6. A 75.0-g sample of a metal is placed in boiling water until its temperature is 100.0°C. A calorimeter contains 100.00 g of water at a temperature of 24.4°C. The metal sample is removed from the boiling water and immediately placed in water in the calorimeter. The final temperature of the metal and water in the calorimeter is 34.9°C. Assuming that the calorimeter provides perfect insulation, what is the specific heat of the metal?

Section 16-3

7. Use **Table 16-6** to determine how much heat is released when 1.00 mole of gaseous methanol condenses to a liquid.

8. Use **Table 16-6** to determine how much heat must be supplied to melt 4.60 grams of ethanol.

Section 16-4

9. Calculate ΔH_{rxn} for the reaction $2C(s) + 2H_2(g) \rightarrow C_2H_4(g)$ given the following thermochemical equations:
$2CO_2(g) + 2H_2O(l) \rightarrow C_2H_4(g) + 3O_2(g)\ \Delta H = 1411\text{ kJ}$
$C(s) + O_2(g) \rightarrow CO_2(g)\ \Delta H = -393.5\text{ kJ}$
$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)\ \Delta H = -572\text{ kJ}$

10. Calculate ΔH_{rxn} for the reaction $HCl(g) + NH_3(g) \rightarrow NH_4Cl(s)$ given the following thermochemical equations:
$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$ $\Delta H = -184$ kJ
$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ $\Delta H = -92$ kJ
$N_2(g) + 4H_2(g) + Cl_2(g) \rightarrow 2NH_4Cl(s)$ $\Delta H = -628$ kJ

Use standard enthalpies of formation from Table 16-7 and Appendix C to calculate $\Delta H°_{rxn}$ for each of the following reactions.

11. $2HF(g) \rightarrow H_2(g) + F_2(g)$
12. $2H_2S(g) + 3O_2(g) \rightarrow 2H_2O(l) + 2SO_2(g)$

Section 16-5 **Predict the sign of ΔS_{system} for each reaction or process.**

13. $FeS(s) \rightarrow Fe^{2+}(aq) + S^{2-}(aq)$
14. $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

Determine if each of the following processes or reactions is spontaneous or nonspontaneous.

15. $\Delta H_{system} = 15.6$ kJ, T = 415 K, $\Delta S_{system} = 45$ J/K
16. $\Delta H_{system} = 35.6$ kJ, T = 415 K, $\Delta S_{system} = 45$ J/K

Chapter 17

Section 17-1

1. In the reaction $A \rightarrow 2B$, suppose that [A] changes from 1.20 mol/L at time = 0 to 0.60 mol/L at time = 3.00 min and that [B] = 0.00 mol/L at time = 0.
a. What is the average rate at which A is consumed in mol/(L·min)?
b. What is the average rate at which B is produced in mol/(L·min)?

Section 17-3

2. What are the overall reaction orders in practice problems 16-18 on page 545?

3. If halving [A] in the reaction $A \rightarrow B$ causes the initial rate to decrease to one fourth its original value, what is the probable rate law for the reaction?

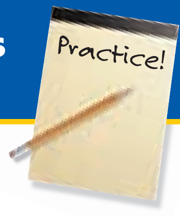
4. Use the data below and the method of initial rates to determine the rate law for the reaction $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

Formation of NO_2 Data

Trial	Initial [NO] (*M*)	Initial [O_2] (*M*)	Initial rate (mol/(L s))
1	0.030	0.020	0.0041
2	0.060	0.020	0.0164
3	0.030	0.040	0.0082

Section 17-4

5. The rate law for the reaction in which one mole of cyclobutane (C_4H_8) decomposes to two moles of ethylene (C_2H_4) at 1273 K is Rate = (87 s^{-1})[C_4H_8]. What is the instantaneous rate of this reaction when
a. [C_4H_8] = 0.0100 mol/L?
b. [C_4H_8] = 0.200 mol/L?



Chapter 18

Section 18-1 **Write equilibrium constant expressions for the following equilibria.**

1. $N_2(g) + O_2(g) \rightleftharpoons 2NO$
2. $3O_2(g) \rightleftharpoons 2O_3(g)$
3. $P_4(g) + 6H_2(g) \rightleftharpoons 4PH_3(g)$
4. $CCl_4(g) + HF(g) \rightleftharpoons CFCl_2(g) + HCl(g)$

Write equilibrium constant expressions for the following equilibria.

5. $NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$
6. $SO_3(g) + H_2O(l) \rightleftharpoons H_2SO_4(l)$

Calculate K_{eq} for the following equilibria.

7. $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$
$[H_2] = 0.0109$, $[I_2] = 0.00290$, $[HI] = 0.0460$
8. $I_2(s) \rightleftharpoons I_2(g)$
$[I_2(g)] = 0.0665$

Section 18-3

9. At a certain temperature, $K_{eq} = 0.0211$ for the equilibrium $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$.
a. What is $[Cl_2]$ in an equilibrium mixture containing 0.865 mol/L PCl_5 and 0.135 mol/L PCl_3?
b. What is $[PCl_5]$ in an equilibrium mixture containing 0.100 mol/L PCl_3 and 0.200 mol/L Cl_2?

10. Use the K_{sp} value for zinc carbonate given in **Table 18-3** to calculate its molar solubility at 298 K.

11. Use the K_{sp} value for iron(II) hydroxide given in **Table 18-3** to calculate its molar solubility at 298 K.

12. Use the K_{sp} value for silver carbonate given in **Table 18-3** to calculate $[Ag^+]$ in a saturated solution at 298 K.

13. Use the K_{sp} value for calcium phosphate given in **Table 18-3** to calculate $[Ca^{2+}]$ in a saturated solution at 298 K.

14. Does a precipitate form when equal volumes of $0.0040M$ $MgCl_2$ and $0.0020M$ K_2CO_3 are mixed? If so, identify the precipitate.

15. Does a precipitate form when equal volumes of $1.2 \times 10^{-4}M$ $AlCl_3$ and $2.0 \times 10^{-3}M$ NaOH are mixed? If so, identify the precipitate.

Chapter 19

Section 19-1

1. Write the balanced formula equation for the reaction between zinc and nitric acid.

2. Write the balanced formula equation for the reaction between magnesium carbonate and sulfuric acid.

3. Identify the base in the reaction
$H_2O(l) + CH_3NH_2(aq) \rightarrow OH^-(aq) + CH_3NH_3^+(aq)$

4. Identify the conjugate base described in the reaction in practice problems 1 and 2.

5. Write the steps in the complete ionization of hydrosulfuric acid.

6. Write the steps in the complete ionization of carbonic acid.

Section 19-2

7. Write the acid ionization equation and ionization constant expression for formic acid (HCOOH).

8. Write the acid ionization equation and ionization constant expression for the hydrogen carbonate ion (HCO^{3-}).

9. Write the base ionization constant expression for ammonia.

10. Write the base ionization expression for aniline ($C_6H_5NH_2$).

Section 19-3

11. Is a solution in which $[H^+] = 1.0 \times 10^{-5}M$ acidic, basic, or neutral?

12. Is a solution in which $[OH^-] = 1.0 \times 10^{-11}M$ acidic, basic, or neutral?

13. What is the pH of a solution in which $[H^+] = 4.5 \times 10^{-4}M$?

14. Calculate the pH and pOH of a solution in which $[OH^-] = 8.8 \times 10^{-3}M$.

15. Calculate the pH and pOH of a solution in which $[H^+] = 2.7 \times 10^{-6}M$.

16. What is $[H^+]$ in a solution having a pH of 2.92?

17. What is $[OH^-]$ in a solution having a pH of 13.56?

18. What is the pH of a 0.000 67*M* H_2SO_4 solution?

19. What is the pH of a 0.000 034*M* NaOH solution?

20. The pH of a 0.200*M* HBrO solution is 4.67. What is the acid's K_a?

21. The pH of a 0.030*M* C_2H_5COOH solution is 3.20. What is the acid's K_a?

Section 19-4

22. Write the formula equation for the reaction between hydroiodic acid and beryllium hydroxide.

23. Write the formula equation for the reaction between perchloric acid and lithium hydroxide.

24. In a titration, 15.73 mL of 0.2346*M* HI solution neutralizes 20.00 mL of a LiOH solution. What is the molarity of the LiOH?

25. What is the molarity of a caustic soda (NaOH) solution if 35.00 mL of solution is neutralized by 68.30 mL of 1.250*M* HCl?

26. Write the chemical equation for the hydrolysis reaction that occurs when sodium hydrogen carbonate is dissolved in water. Is the resulting solution acidic, basic, or neutral?

27. Write the chemical equation for any hydrolysis reaction that occurs when cesium chloride is dissolved in water. Is the resulting solution acidic, basic, or neutral?

Chapter 20

Section 20-1 **Identify the following information for each problem. What element is oxidized? Reduced? What is the oxidizing agent? Reducing agent?**

1. $2P + 3Cl_2 \rightarrow 2PCl_3$

2. $C + H_2O \rightarrow CO + H_2$

3. Determine the oxidation number for each element in the following compounds.
   a. Na_2SeO_3
   b. $HAuCl_4$
   c. H_3BO_3

4. Determine the oxidation number for the following compounds or ions.
   a. P_4O_8
   b. Na_2O_2 (hint: this is like H_2O_2)
   c. AsO_4^{-3}

Section 20-2

5. How many electrons will be lost or gained in each of the following half-reactions? Identify whether it is an oxidation or reduction.
   a. $Cr \rightarrow Cr^{3+}$
   b. $O_2 \rightarrow O^{2-}$
   c. $Fe^{+2} \rightarrow Fe^{3+}$

6. Balance the following reaction by the oxidation number method: $MnO_4^- + CH_3OH \rightarrow MnO_2 + HCHO$ (acidic). (Hint: assign the oxidation of hydrogen and oxygen as usual and solve for the oxidation number of carbon.)

7. Balance the following reaction by the oxidation number method: $Zn + HNO_3 \rightarrow ZnO + NO_2 + NH_3$

8. Use the oxidation number method to balance these net ionic equations:
   a. $SeO_3^{2-} + I^- \rightarrow Se + I_2$ (acidic solution)
   b. $NiO_2 + S_2O_3^{2-} \rightarrow Ni(OH)_2 + SO_3^{2-}$ (acidic solution)

Section 20-3 **Use the half-reaction method to balance the following redox equations.**

9. $Zn(s) + HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$
10. $MnO_4^-(aq) + H_2SO_3(aq) \longrightarrow Mn^{2+}(aq) + HSO_4^-(aq) + H_2O(l)$ (acidic solution)
11. $NO_2(aq) + OH^-(aq) \longrightarrow NO_2^-(aq) + NO_3^-(aq) + H_2O(l)$ (basic solution)
12. $HS^-(aq) + IO_3^-(aq) \longrightarrow I^-(aq) + S(s) + H_2O(l)$ (acidic solution)

Chapter 21

Section 21-1

1. Calculate the cell potential for each of the following.
   a. $Co^{2+}(aq) + Al(s) \rightarrow Co(s) + Al^{3+}(aq)$
   b. $Hg^{2+}(aq) + Cu(s) \rightarrow Cu^{2+}(aq) + Hg(s)$
   c. $Zn(s) + Br_2(l) \rightarrow Br^{1-}(aq) + Zn^{2+}(aq)$

2. Calculate the cell potential to determine whether the reaction will occur spontaneously or not spontaneously. For each reaction that is not spontaneous, correct the reactants or products so that a reaction would occur spontaneously.
   a. $Ni^{2+}(aq) + Al(s) \rightarrow Ni(s) + Al^{3+}(aq)$
   b. $Ag^+(aq) + H_2(g) \rightarrow Ag(s) + H^+(aq)$
   c. $Fe^{2+}(aq) + Cu(s) \rightarrow Fe(s) + Cu^{2+}(aq)$

Chapter 22

Section 22-1

1. Draw the structure of the following branched alkanes.
a. 2,2,4-trimethylheptane
b. 4-isopropyl-2-methylnonane

Section 22-2

2. Draw the structure of each of the following cycloalkanes.
a. 1-ethyl-2-methylcyclobutane
b. 1,3-dibutylcyclohexane

Section 22-3

3. Draw the structure of each of the following alkenes.
a. 1,4-hexadiene
b. 2,3-dimethyl-2-butene
c. 4-propyl-1-octene
d. 2,3-diethylcyclohexene

Chapter 23

Section 23-1

1. Draw the structures of the following alkyl halides.
a. chloroethane
b. chloromethane
c. 1-fluoropentane
d. 1,3-dibromocyclohexane
e. 1,2-dibromo-3-chloropropane

Chapter 25

Section 25-2

1. Write balanced equations for each of the following decay processes.
a. Alpha emission of $^{244}_{96}Cm$
b. Positron emission of $^{70}_{33}As$
c. Beta emission of $^{210}_{83}Bi$
d. Electron capture by $^{116}_{51}Sb$

2. $^{47}_{20}Ca \rightarrow \,^{0}_{-1}\beta + ?$

3. $^{240}_{95}Am + ? \rightarrow \,^{243}_{97}Bk + \,^{1}_{0}n$

Section 25-3

4. How much time has passed if 1/8 of an original sample of radon-222 is left? Use **Table 25-5** for half-life information.

5. If a basement air sample contains 3.64 μg of radon-222, how much radon will remain after 19 days?

6. Cobalt-60, with a half-life of 5 years, is used in cancer radiation treatments. If a hospital purchases a supply of 30.0 g, how much would be left after 15 years?