3 Stoichiometry: Calculations with Chemical Formulas and Equations

Visualizing Concepts

3.1 Reactant A = red, reactant B = blue

Overall, 4 red A_2 molecules + 4 blue B atoms → 4 A_2B molecules

Since 4 is a common factor, this equation reduces to equation (a).

3.3 (a) There are twice as many O atoms as N atoms, so the empirical formula of the original compound is NO_2.

(b) No, because we have no way of knowing whether the empirical and molecular formulas are the same. NO_2 represents the simplest ratio of atoms in a molecule but not the only possible molecular formula.

3.5 (a) *Analyze*. Given the molecular model, write the molecular formula.

Plan. Use the colors of the atoms (spheres) in the model to determine the number of atoms of each element.

Solve. Observe 2 gray C atoms, 5 white H atoms, 1 blue N atom, 2 red O atoms. $C_2H_5NO_2$

(b) *Plan*. Follow the method in Sample Exercise 3.9. Calculate formula weight in amu and molar mass in grams.

2 C atoms = 2(12.0 amu) = 24.0 amu

5 H atoms = 5(1.0 amu) = 5.0 amu

1 N atoms = 1(14.0 amu) = 14.0 amu

2 O atoms = 2(16.0 amu) = <u>32.0 amu</u>

75.0 amu

Formula weight = 75.0 amu, molar mass = 75.0 g/mol

(c) *Plan*. The molar mass of a substance provides the factor for converting moles to grams (or grams to moles).

Solve. 3 mol glycine $\times \dfrac{75.0 \text{ g glycine}}{\text{mol}} = 225 \text{ g glycine}$

(d) *Plan*. Use the definition of mass % and the results from parts (a) and (b) above to find mass % N in glycine.

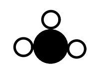
Solve. $$\text{mass\% N} = \frac{\text{g N}}{\text{g } C_2H_5NO_2} \times 100$$

Assume 1 mol $C_2H_5NO_2$. From the molecular formula of glycine [part (a)], there is 1 mol N/mol glycine.

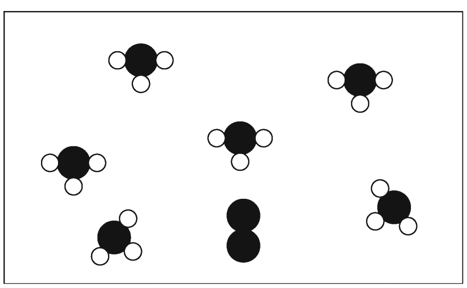
$$\text{mass\% N} = \frac{1 \times (\text{molar mass N})}{\text{molar mass glycine}} \times 100 = \frac{14.0 \text{ g}}{75.0 \text{ g}} \times 100 = 18.7\%$$

3.7 *Analyze*. Given a box diagram and formulas of reactants, draw a box diagram of products.

Plan. Write and balance the chemical equation. Determine combining ratios of elements and decide on limiting reactant. Draw a box diagram of products, containing the correct number of product molecules and only excess reactant.

Solve. $N_2 + 3H_2 \longrightarrow 3NH_3$. N_2 =  , NH_3 = 

Each N atom (1/2 of an N_2 molecule) reacts with 3 H atoms (1.5 H_2 molecules) to form an NH_3 molecule. Eight N atoms (4 N_2 molecules) require 24 H atoms (12 H_2 molecules) for complete reaction. Only 9 H_2 molecules are available, so H_2 is the limiting reactant. Nine H_2 molecules (18 H atoms) determine that 6 NH_3 molecules are produced. One N_2 molecule is in excess.



Check. Verify that mass is conserved in your solution, that the number and kinds of atoms are the same in reactant and product diagrams. In this example, there are 8 N atoms and 18 H atoms in both diagrams, so mass is conserved.

Balancing Chemical Equations

3.9 (a) In balancing chemical equations, the law of conservation of mass, that atoms are neither created nor destroyed during the course of a reaction, is observed. This means that the **number** and **kinds** of atoms on both sides of the chemical equation must be the same.

(b) Subscripts in chemical formulas should not be changed when balancing equations, because changing the subscript changes the identity of the compound (law of constant composition).

(c) liquid water = $H_2O(l)$; water vapor = $H_2O(g)$; aqueous sodium chloride =

NaCl(aq); solid sodium chloride = NaCl(s)

3.10 (a) In a CO molecule, there is one O atom bound to C. 2CO indicates that there are **two CO molecules**, each of which contains one C and one O atom. Adding a subscript 2 to CO to form CO_2 means that there are **two O atoms** bound to one C in a CO_2 molecule. The composition of the different molecules, CO_2 and CO, is different and the physical and chemical properties of the two compounds they constitute are very different. The subscript 2 changes molecular composition and thus properties of the compound. The prefix 2 indicates how many molecules (or moles) of the original compound are under consideration.

(b) Yes. There are the same number and kinds of atoms on the reactants side and the products side of the equation.

3.11 (a) $\mathbf{2}CO(g) + O_2(g) \rightarrow \mathbf{2}CO_2(g)$

(b) $N_2O_5(g) + H_2O(l) \rightarrow \mathbf{2}HNO_3(aq)$

(c) $CH_4(g) + \mathbf{4}Cl_2(g) \rightarrow CCl_4(l) + \mathbf{4}HCl(g)$

(d) $Al_4C_3(s) + \mathbf{12}H_2O(l) \rightarrow \mathbf{4}Al(OH)_3(s) + \mathbf{3}CH_4(g)$

(e) $\mathbf{2}C_5H_{10}O_2(l) + \mathbf{13}O_2(g) \rightarrow \mathbf{10}CO_2(g) + \mathbf{10}H_2O(l)$

(f) $\mathbf{2}Fe(OH)_3(s) + \mathbf{3}H_2SO_4(aq) \rightarrow Fe_2(SO_4)_3(aq) + \mathbf{6}H_2O(l)$

(g) $Mg_3N_2(s) + \mathbf{4}H_2SO_4(aq) \rightarrow \mathbf{3}MgSO_4(aq) + (NH_4)_2SO_4(aq)$

3.13 (a) $CaC_2(s) + \mathbf{2}H_2O(l) \rightarrow Ca(OH)_2(aq) + C_2H_2(g)$

(b) $2KClO_3(s) \xrightarrow{\Delta} 2KCl(s) + 3O_2(g)$

(c) $Zn(s) + H_2SO_4(aq) \rightarrow H_2(g) + ZnSO_4(aq)$

(d) $PCl_3(l) + \mathbf{3}H_2O(l) \rightarrow H_3PO_3(aq) + \mathbf{3}HCl(aq)$

(e) $\mathbf{3}H_2S(g) + \mathbf{2}Fe(OH)_3(s) \rightarrow Fe_2S_3(s) + \mathbf{6}H_2O(g)$

Patterns of Chemical Reactivity

3.15 (a) When a metal reacts with a nonmetal, an ionic compound forms. The combining ratio of the atoms is such that the total positive charge on the metal cation(s) is equal to the total negative charge on the nonmetal anion(s). Determine the formula by balancing the positive and negative charges in the ionic product. All ionic compounds are solids. $2\ Na(s) + Br_2(l) \rightarrow 2NaBr(s)$

(b) The second reactant is oxygen gas from the air, $O_2(g)$. The products are $CO_2(g)$ and $H_2O(l)$. $2C_6H_6(l) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(l)$.

3.17 (a) $Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$

(b) $BaCO_3(s) \xrightarrow{\Delta} BaO(s) + CO_2(g)$

(c) $C_8H_8(l) + \mathbf{10}O_2(g) \rightarrow \mathbf{8}CO_2(g) + \mathbf{4}H_2O(l)$

(d) CH_3OCH_3 is C_2H_6O. $C_2H_6O(l) + \mathbf{3}O_2(g) \rightarrow \mathbf{2}CO_2(g) + \mathbf{3}H_2O(l)$

3.19 (a) $\mathbf{2}Al(s) + \mathbf{3}Cl_2(g) \rightarrow \mathbf{2}AlCl_3(s)$ combination

(b) $C_2H_4(g) + \mathbf{3}O_2(g) \rightarrow \mathbf{2}CO_2(g) + \mathbf{2}H_2O(l)$ combustion

(c) $\mathbf{6}Li(s) + N_2(g) \rightarrow \mathbf{2}Li_3N(s)$ combination

(d) $PbCO_3(s) \rightarrow PbO(s) + CO_2(g)$ decomposition

(e) $C_7H_8O_2(l) + \mathbf{8}O_2(g) \rightarrow \mathbf{7}CO_2(g) + \mathbf{4}H_2O(l)$ combustion

Formula Weights

3.21 *Analyze*. Given molecular formula or name, calculate formula weight.

Plan. If a name is given, write the correct molecular formula. Then, follow the method in Sample Exercise 3.5. *Solve*.

(a) HNO_3: 1(1.0) + 1(14.0) + 3(16.0) = 63.0 amu

(b) $KMnO_4$: 1(39.1) + 1(54.9) + 4(16.0) = 158.0 amu

(c) $Ca_3(PO_4)_2$: 3(40.1) + 2(31.0) + 8(16.0) = 310.3 amu

(d) SiO_2: 1(28.1) + 2(16.0) = 60.1 amu

(e) Ga_2S_3: 2(69.7) + 3(32.1) = 235.7 amu

(f) $Cr_2(SO_4)_3$: 2(52.0) + 3(32.1) + 12(16.0) = 392.3 amu

(g) PCl_3: 1(31.0) + 3(35.5) = 137.5 amu

3.23 *Plan*. Calculate the formula weight (FW), then the mass % oxygen in the compound. *Solve*.

(a) $C_{17}H_{19}NO_3$: FW = 17(12.0) + 19(1.0) + 1(14.0) + 3(16.0) = 285.0 amu

$$\%O = \frac{3(16.0)\text{ amu}}{285.0\text{ amu}} \times 100 = 16.842 = 16.8\%$$

(b) $C_{18}H_{21}NO_3$: FW = 18(12.0) + 21(1.0) + 1(14.0) + 3(16.0) = 299.0 amu

$$\%O = \frac{3(16.0)\text{ amu}}{299.0\text{ amu}} \times 100 = 16.054 = 16.1\%$$

(c) $C_{17}H_{21}NO_4$: FW = 17(12.0) + 21(1.0) + 1(14.0) + 4(16.0) = 303.0 amu

$$\%O = \frac{4(16.0)\text{ amu}}{303.0\text{ amu}} \times 100 = 21.122 = 21.1\%$$

(d) $C_{22}H_{24}N_2O_8$: FW = 22(12.0) + 24(1.0) + 2(14.0) + 8(16.0) = 444.0 amu

$$\%O = \frac{8(16.0)\text{ amu}}{444.0\text{ amu}} \times 100 = 28.829 = 28.8\%$$

(e) $C_{41}H_{64}O_{13}$: FW = 4(12.0) + 64(1.0) + 13(16.0) = 764.0 amu

$$\%O = \frac{13(16.0)\text{ amu}}{764\text{ amu}} \times 100 = 27.225 = 27.2\%$$

(f) $C_{66}H_{75}Cl_2N_9O_{24}$: FW = 66(12.0)+75(1.0)+2(35.5)+9(14.0)+24(16.0) = 1448.0 amu

$$\%O = \frac{24(16.0)\text{ amu}}{1448.0\text{ amu}} \times 100 = 26.519 = 26.5\%$$

3.25 *Plan*. Follow the logic for calculating mass % C given in Sample Exercise 3.6. *Solve*.

(a) C_7H_6O: FW = 7(12.0) + 6(1.0) + 1(16.0) = 106.0 amu

$$\%C = \frac{7(12.0)\text{ amu}}{106.0\text{ amu}} \times 100 = 79.2\%$$

(b) $C_8H_8O_3$: FW = 8(12.0) + 8(1.0) + 3(16.0) = 152.0 amu

$$\%C = \frac{8(12.0)\text{ amu}}{152.0\text{ amu}} \times 100 = 63.2\%$$

(c) $C_7H_{14}O_2$: FW = 7(12.0) + 14(1.0) + 2(16.0) = 130.0 amu

$$\%C = \frac{7(12.0)\text{ amu}}{130.0\text{ amu}} \times 100 = 64.6\%$$

Avogadro's Number and the Mole

3.27 (a) 6.022×10^{23}. This is the number of objects in a mole of anything.

(b) The formula weight of a substance in amu has the same numerical value as the molar mass expressed in grams.

3.29 *Plan*. Since the mole is a counting unit, use it as a basis of comparison; determine the total moles of atoms in each given quantity. *Solve*.

23 g Na contains 1 mol of atoms

0.5 mol H_2O contains (3 atoms × 0.5 mol) = 1.5 mol atoms

6.0×10^{23} N_2 molecules contains (2 atoms × 1 mol) = 2 mol atoms

3.31 *Analyze*. Given: 160 lb/person; Avogadro's number of people, 6.022×10^{23} people. Find: mass in kg of Avogadro's number of people; compare with mass of Earth.

Plan. people → mass in lb → mass in kg; mass of people / mass of Earth

$$\textit{Solve}.\ 6.022 \times 10^{23}\text{ people} \times \frac{160\text{ lb}}{\text{person}} \times \frac{1\text{ kg}}{2.2046\text{ lb}} = 4.370 \times 10^{25} = 4.37 \times 10^{25}\text{ or }4.4 \times 10^{25}\text{ kg}$$

$$\frac{4.370 \times 10^{25}\text{ kg of people}}{5.98 \times 10^{24}\text{ kg Earth}} = 7.31\text{ or }7.3$$

One mole of people weighs 7.31 times as much as Earth.

Check. This mass of people is reasonable since Avogadro's number is large.

Estimate: 160 lb ≈ 70 kg; $6 \times 10^{23} \times 70 = 420 \times 10^{23} = 4.2 \times 10^{25}$ kg

3.33 (a) *Analyze*. Given: 0.105 mol sucrose, $C_{12}H_{22}O_{11}$. Find: mass in g.

Plan. Use molar mass (g/mol) of $C_{12}H_{22}O_{11}$ to find g $C_{12}H_{22}O_{11}$.

Solve. molar mass = 12(12.0107) + 22(1.00794) + 11(15.9994) = 342.296 = 342.30

$$0.105 \text{ mol } C_{12}H_{22}O_{11} \times \frac{342.30 \text{ g}}{1 \text{ mol}} = 35.942 = 35.9 \text{ g } C_{12}H_{22}O_{11}$$

Check. 0.1(342) = 34.2 g. The calculated result is reasonable.

(b) *Analyze*. Given: mass. Find: moles. *Plan*. Use molar mass of $Zn(NO_3)_2$.

Solve. molar mass = 1(65.39) + 2(14.01) + 6(16.00) = 189.41 = 189.4

$$143.5 \text{ g } Zn(NO_3)_2 \times \frac{1 \text{ mol}}{189.41 \text{ g}} = 0.7576 \text{ mol } Zn(NO_3)_2$$

Check. 140/180 ≈ 7/9 = 0.78 mol

(c) *Analyze*. Given: moles. Find: molecules. *Plan*. Use Avogadro's number.

$$\textit{Solve}\; 1.0 \times 10^{-6} \text{ mol } CH_3CH_2OH \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 6.022 \times 10^{17}$$

$$= 6.0 \times 10^{17}\ CH_3CH_2OH \text{ molecules}$$

Check. $(1.0 \times 10^{-6})(6 \times 10^{23}) = 6 \times 10^{17}$

(d) *Analyze*. Given: mol NH_3. Find: N atoms.

Plan. mol NH_3 → mol N atoms → N atoms

$$\textit{Solve}\; 0.410 \text{ mol } NH_3 \times \frac{1 \text{ mol N atoms}}{1 \text{ mol } NH_3} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}}$$

$$= 2.47 \times 10^{23} \text{ C atoms}$$

Check. $(0.4)(6 \times 10^{23}) = 2.4 \times 10^{23}$.

3.35 *Analyze/Plan*. See Solution 3.33 for stepwise problem-solving approach. *Solve*.

(a) $(NH_4)_3PO_4$ molar mass = 3(14.007) + 12(1.008) + 1(30.974) + 4(16.00) = 149.091

= 149.1 g/mol

$$2.50 \times 10^{-3} \text{ mol } (NH_4)_3PO_4 \times \frac{149.1 \text{ g } (NH_4)_3PO_4}{1 \text{ mol}} = 0.373 \text{ g } (NH_4)_3PO_4$$

(b) $AlCl_3$ molar mass = 26.982 + 3(35.453) = 133.341 = 133.34 g/mol

$$0.2550 \text{ g } AlCl_3 \times \frac{1 \text{ mol}}{133.34 \text{ g } AlCl_3} \times \frac{3 \text{ mol } Cl^-}{1 \text{ mol } AlCl_3} = 5.737 \times 10^{-3} \text{ mol } Cl^-$$

(c) $C_8H_{10}N_4O_2$ molar mass = 8(12.01) + 10(1.008) + 4(14.01) + 2(16.00) = 194.20

= 194.2 g/mol

$$7.70 \times 10^{20} \text{ molecules} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} \times \frac{194.2 \text{ g } C_8H_{10}N_4O_2}{1 \text{ mol caffeine}}$$

$= 0.248$ g $C_8H_{10}N_4O_2$

(d) $\dfrac{0.406 \text{ g cholesterol}}{0.00105 \text{ mol}} = 387 \text{ g cholesterol/mol}$

3.37 (a) $C_6H_{10}OS_2$ molar mass = 6(12.01) + 10(1.008) + 1(16.00) + 2(32.07) = 162.28

= 162.3 g/mol

(b) *Plan*. mg → g → mol *Solve*.

$$5.00 \text{ mg allicin} \times \frac{1 \times 10^{-3} \text{ g}}{1 \text{ mg}} \times \frac{1 \text{ mol}}{162.3 \text{ g}} = 3.081 \times 10^{-5} = 3.08 \times 10^{-5} \text{ mol allicin}$$

Check. 5.00 mg is a small mass, so the small answer is reasonable.

$(5 \times 10^{-3})/200 = 2.5 \times 10^{-5}$

(c) *Plan*. Use mol from part (b) and Avogadro's number to calculate molecules.

$$\text{Solve. } 3.081 \times 10^{-5} \text{ mol allicin} \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} = 1.855 \times 10^{19}$$

$$= 1.86 \times 10^{19} \text{ allicin molecules}$$

Check. $(3 \times 10^{-5})(6 \times 10^{23}) = 18 \times 10^{18} = 1.8 \times 10^{19}$

(d) *Plan*. Use molecules from part (c) and molecular formula to calculate S atoms.

$$\textit{Solve. } 1.855 \times 10^{19} \text{ allicin molecules} \times \frac{2 \text{ S atoms}}{1 \text{ allicin molecule}} = 3.71 \times 10^{19} \text{ S atoms}$$

Check. Obvious.

3.39 (a) *Analyze*. Given: $C_6H_{12}O_6$, 1.250×10^{21} C atoms. Find: H atoms.

Plan. Use molecular formula to determine number of H atoms that are present with 1.250×10^{21} C atoms. *Solve*.

$$\frac{12 \text{ H atoms}}{6 \text{ C atoms}} = \frac{2 \text{ H}}{1 \text{ C}} \times 1.250 \times 10^{21} \text{ C atoms} = 2.500 \times 10^{21} \text{ H atoms}$$

Check. $(2 \times 1 \times 10^{21}) = 2 \times 10^{21}$

(b) *Plan*. Use molecular formula to find the number of glucose molecules that contain 1.250×10^{21} C atoms. *Solve*.

$$\frac{1 \text{ } C_6H_{12}O_6 \text{ molecule}}{6 \text{ C atoms}} \times 1.250 \times 10^{21} \text{ C atoms} = 2.0833 \times 10^{20}$$

$$= 2.083 \times 10^{20} \text{ } C_6H_{12}O_6 \text{ molecules}$$

Check. $(12 \times 10^{20}/6) = 2 \times 10^{20}$

(c) *Plan*. Use Avogadro's number to change molecules → mol. *Solve*.

$$2.0833 \times 10^{20} \text{ } C_6H_{12}O_6 \text{ molecules} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}}$$

$$= 3.4595 \times 10^{-4} = 3.460 \times 10^{-4} \text{ mol } C_6H_{12}O_6$$

Check. $(2 \times 10^{20})/(6 \times 10^{23}) = 0.33 \times 10^{-3} = 3.3 \times 10^{-4}$

(d) *Plan*. Use molar mass to change mol → g. *Solve*.

1 mole of $C_6H_{12}O_6$ weighs 180.0 g (Sample Exercise 3.9)

$$3.4595 \times 10^{-4} \text{ mol } C_6H_{12}O_6 \times \frac{180.0 \text{ g } C_6H_{12}O_6}{1 \text{ mol}} = 0.06227 \text{ g } C_6H_{12}O_6$$

Check. $3.5 \times 180 = 630$; $630 \times 10^{-4} = 0.063$

3.41 *Analyze*. Given: g C_2H_3Cl/L. Find: mol/L, molecules/L.

Plan. The /L is constant throughout the problem, so we can ignore it. Use molar mass for g → mol, Avogadro's number for mol → molecules. *Solve*.

$$\frac{2.0 \times 10^{-6} \text{ g } C_2H_3Cl}{1 \text{ L}} \times \frac{1 \text{ mol } C_2H_3Cl}{62.50 \text{ g } C_2H_3Cl} = 3.20 \times 10^{-8} = 3.2 \times 10^{-8} \text{ mol } C_2H_3Cl/L$$

$$\frac{3.20 \times 10^{-8} \text{ mol } C_2H_3Cl}{1 \text{ L}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 1.9 \times 10^{16} \text{ molecules/L}$$

Check. $(200 \times 10^{-8})/60 = 2.5 \times 10^{-8}$ mol

$(2.5 \times 10^{-8}) \times (6 \times 10^{23}) = 15 \times 10^{15} = 1.5 \times 10^{16}$

Empirical Formulas

3.43 (a) *Analyze*. Given: moles. Find: empirical formula.

Plan. Find the **simplest ratio of moles** by dividing by the smallest number of moles present.

Solve. 0.0130 mol C / 0.0065 = 2

0.039 mol H / 0.0065 = 6

0.0065 mol O / 0.0065 = 1

The empirical formula is C_2H_6O.

Check. The subscripts are simple integers.

(b) *Analyze*. Given: grams. Find: empirical formula.

Plan. Calculate the moles of each element present, then the simplest ratio of moles.

Solve. $$11.66 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 0.2088 \text{ mol Fe};\ 0.2088/0.2088 = 1$$

$$5.01 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.3131 \text{ mol O};\ 0.3131/0.2088 = 1.5$$

Multiplying by two, the integer ratio is 2 Fe : 3 O; the empirical formula is Fe_2O_3.

Check. The subscripts are simple integers.

(c) *Analyze*. Given: mass %. Find: empirical formulas.

Plan. Assume 100 g sample, calculate moles of each element, find the simplest ratio of moles.

$$Solve.\ 40.0\ \text{g C} \times \frac{1\ \text{mol C}}{12.01\ \text{g C}} = 3.33\ \text{mol C};\ 3.33/3.33 = 1$$

$$6.7\ \text{g H} \times \frac{1\ \text{mol H}}{1.008\ \text{mol H}} = 6.65\ \text{mol H};\ 6.65/3.33 \approx 2$$

$$53.3\ \text{g O} \times \frac{1\ \text{mol O}}{16.00\ \text{mol O}} = 3.33\ \text{mol O};\ 3.33/3.33 = 1$$

The empirical formula is CH_2O.

Check. The subscripts are simple integers.

3.45 *Analyze/Plan*. The procedure in all these cases is to assume 100 g of sample, calculate the number of moles of each element present in that 100 g, then obtain the ratio of moles as smallest whole numbers. *Solve*.

(a) $$10.4\ \text{g C} \times \frac{1\ \text{mol C}}{12.01\ \text{g C}} = 0.866\ \text{mol C};\ 0.866/0.866 = 1$$

$$27.8\ \text{g S} \times \frac{1\ \text{mol S}}{32.07\ \text{g S}} = 0.867\ \text{mol S};\ 0.867/0.866 \approx 1$$

$$61.7\ \text{g Cl} \times \frac{1\ \text{mol Cl}}{35.45\ \text{g Cl}} = 1.74\ \text{mol Cl};\ 1.74/0.866 \approx 2$$

The empirical formula is $CSCl_2$.

(b) $$21.7\ \text{g C} \times \frac{1\ \text{mol C}}{12.01\ \text{g C}} = 1.81\ \text{mol C};\ 1.81/0.600 \approx 3$$

$$9.6\ \text{g O} \times \frac{1\ \text{mol O}}{16.00\ \text{g O}} = 0.600\ \text{mol O};\ 0.600/0.600 = 1$$

$$68.7\ \text{g F} \times \frac{1\ \text{mol F}}{19.00\ \text{g F}} = 3.62\ \text{mol F};\ 3.62/0.600 \approx 6$$

The empirical formula is C_3OF_6.

(c) $$32.79\ \text{g Na} \times \frac{1\ \text{mol Na}}{22.99\ \text{g Na}} = 1.426\ \text{mol Na};\ 1.426/0.4826 \approx 3$$

$$13.02\ \text{g Al} \times \frac{1\ \text{mol Al}}{26.98\ \text{g Al}} = 0.4826\ \text{mol Al};\ 0.4826/0.4826 = 1$$

$$54.19\ \text{g F} \times \frac{1\ \text{mol F}}{19.00\ \text{g F}} = 2.852\ \text{mol F};\ 2.852/0.4826 \approx 6$$

The empirical formula is Na_3AlF_6.

3.47 *Analyze*. Given: empirical formula, molar mass. Find: molecular formula.

Plan. Calculate the empirical formula weight (FW); divide FW by molar mass (MM) to calculate the integer that relates the empirical and molecular formulas. Check. If FW/MM is an integer, the result is reasonable. *Solve*.

(a) $$\text{FW } CH_2 = 12 + 2(1) = 14 \quad \frac{MM}{FW} = \frac{84}{14} = 6$$

The subscripts in the empirical formula are multiplied by 6. The molecular formula is C_6H_{12}.

(b) $\text{FW NH}_2\text{Cl} = 14.01 + 2(1.008) + 35.45 = 51.48. \quad \frac{\text{MM}}{\text{FW}} = \frac{51.5}{51.5} = 1$

The empirical and molecular formulas are NH_2Cl.

3.49 *Analyze*. Given: mass %, molar mass. Find: molecular formula.

Plan. Use the plan detailed in Solution 3.45 to find an empirical formula from mass % data. Then use the plan detailed in 3.47 to find the molecular formula. Note that some indication of molar mass must be given, or the molecular formula cannot be determined. *Check*. If there is an integer ratio of moles and MM/ FW is an integer, the result is reasonable. *Solve*.

(a) $92.3 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 7.685 \text{ mol C}; \; 7.685/7.639 = 1.006 \approx 1$

$7.7 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 7.639 \text{ mol H}; \; 7.639/7.639 = 1$

The empirical formula is CH, FW = 13.

$\frac{\text{MM}}{\text{FW}} = \frac{104}{13} = 8$; the molecular formula is C_8H_8.

(b) $49.5 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.12 \text{ mol C}; \; 4.12/1.03 = 4$

$5.15 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 5.11 \text{ mol H}; \; 5.11/1.03 = 5$

$28.9 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 2.06 \text{ mol N}; \; 2.06/1.03 = 2$

$16.5 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.03 \text{ mol O}; \; 1.03/1.03 = 1$

Thus, $C_4H_5N_2O$, FW = 97. If the molar mass is about 195, a factor of 2 gives the molecular formula $C_8H_{10}N_4O_2$.

(c) $35.51 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 2.96 \text{ mol C}; \; 2.96/0.592 = 5$

$4.77 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.73 \text{ mol H}; \; 4.73/0.592 = 7.99 \approx 8$

$37.85 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.37 \text{ mol O}; \; 2.37/0.592 = 4$

$8.29 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.592 \text{ mol N}; \; 0.592/0.592 = 1$

$13.60 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} = 0.592 \text{ mol Na}; \; 0.592/0.592 = 1$

The empirical formula is $C_5H_8O_4NNa$, FW = 169 g. Since the empirical formula weight and molar mass are approximately equal, the empirical and molecular formulas are both $NaC_5H_8O_4N$.

3.51 (a) *Analyze*. Given: mg CO_2, mg H_2O Find: empirical formula of hydrocarbon, C_xH_y

Plan. Upon combustion, all C → CO_2, all H → H_2O.

mg CO_2 → g CO_2 → mol C; mg H_2O → g H_2O, mol H

Find simplest ratio of moles and empirical formula. *Solve*.

$$5.86 \times 10^{-3}\ \text{g } CO_2 \times \frac{1\ \text{mol } CO_2}{44.01\ \text{g } CO_2} \times \frac{1\ \text{mol C}}{1\ \text{mol } CO_2} = 1.33 \times 10^{-4}\ \text{mol C}$$

$$1.37 \times 10^{-3}\ \text{g } H_2O \times \frac{1\ \text{mol } H_2O}{18.02\ \text{g } H_2O} \times \frac{2\ \text{mol H}}{1\ \text{mol } H_2O} = 1.52 \times 10^{-4}\ \text{mol H}$$

Dividing both values by 1.33×10^{-4} gives C:H of 1:1.14. This is not "close enough" to be considered 1:1. No obvious multipliers (2, 3, 4) produce an integer ratio. Testing other multipliers (trial and error!), the correct factor seems to be 7. The empirical formula is C_7H_8.

Check. See discussion of C:H ratio above.

(b) *Analyze*. Given: g of menthol, g CO_2, g H_2O, molar mass. Find: molecular formula.

Plan/Solve. Calculate mol C and mol H in the sample.

$$0.2829\ \text{g } CO_2 \times \frac{1\ \text{mol } CO_2}{44.01\ \text{g } CO_2} \times \frac{1\ \text{mol C}}{1\ \text{mol } CO_2} = 0.0064281 = 0.006428\ \text{mol C}$$

$$0.1159\ \text{g } H_2O \times \frac{1\ \text{mol } H_2O}{18.02\ \text{g } H_2O} \times \frac{2\ \text{mol H}}{1\ \text{mol } H_2O} = 0.012863 = 0.01286\ \text{mol H}$$

Calculate g C, g H and get g O by subtraction.

$$0.0064281\ \text{mol C} \times \frac{12.01\ \text{g C}}{1\ \text{mol C}} = 0.07720\ \text{g C}$$

$$0.012863\ \text{mol H} \times \frac{1.008\ \text{g H}}{1\ \text{mol H}} = 0.01297\ \text{g H}$$

mass O = 0.1005 g sample – (0.07720 g C + 0.01297 g H) = 0.01033 g O

Calculate mol O and find integer ratio of mol C: mol H: mol O.

$$0.01033\ \text{g O} \times \frac{1\ \text{mol O}}{16.00\ \text{g O}} = 6.456 \times 10^{-4}\ \text{mol O}$$

Divide moles by 6.456×10^{-4}.

$$\text{C}: \frac{0.006428}{6.456 \times 10^{-4}} \approx 10;\quad \text{H}: \frac{0.01286}{6.456 \times 10^{-4}} \approx 20;\quad \text{O}: \frac{6.456 \times 10^{-4}}{6.456 \times 10^{-4}} = 1$$

The empirical formula is $C_{10}H_{20}O$.

$$FW = 10(12) + 20(1) + 16 = 156; \frac{M}{FW} = \frac{156}{156} = 1$$

The molecular formula is the same as the empirical formula, $C_{10}H_{20}O$.

Check. The mass of O wasn't negative or greater than the sample mass; empirical and molecular formulas are reasonable.

3.53 *Analyze*. Given 2.558 g $Na_2CO_3 \cdot xH_2O$, 0.948 g Na_2CO_3. Find: x.

Plan. The reaction involved is $Na_2CO_3 \cdot xH_2O(s) \rightarrow Na_2CO_3(s) + xH_2O(g)$.

Calculate the mass of H_2O lost and then the mole ratio of Na_2CO_3 and H_2O.

Solve. g H_2O lost = 2.558 g sample – 0.948 g Na_2CO_3 = 1.610 g H_2O

$$0.948\ g\ Na_2CO_3 \times \frac{1\ mol\ Na_2CO_3}{106.0\ g\ Na_2CO_3} = 0.00894\ mol\ Na_2CO_3$$

$$1.610\ g\ H_2O \times \frac{1\ mol\ H_2O}{18.02\ g\ H_2O} = 0.08935\ mol\ H_2O$$

The formula is $Na_2CO_3 \cdot$ **<u>10</u>** H_2O.

Check. x is an integer.

Calculations Based on Chemical Equations

3.55 The mole ratios implicit in the coefficients of a balanced chemical equation express the fundamental relationship between amounts of reactants and products. If the equation is not balanced, the mole ratios will be incorrect and lead to erroneous calculated amounts of products.

3.57 $Na_2SiO_3(s) + 8HF(aq) \rightarrow H_2SiF_6(aq) + 2NaF(aq) + 3H_2O(l)$

(a) *Analyze*. Given: mol Na_2SiO_3. Find: mol HF. *Plan*. Use the mole ratio $8HF{:}1Na_2SiO_3$ from the balanced equation to relate moles of the two reactants.

Solve.

$$0.300\ mol\ Na_2SiO_3 \times \frac{8\ mol\ HF}{1\ mol\ Na_2SiO_3} = 2.40\ mol\ HF$$

Check. Mol HF should be greater than mol Na_2SiO_3.

(b) *Analyze*. Given: mol HF. Find: g NaF. *Plan*. Use the mole ratio 2NaF:8HF to change mol HF to mol NaF, then molar mass to get NaF. *Solve*.

$$0.500\ mol\ HF \times \frac{2\ mol\ NaF}{8\ mol\ HF} \times \frac{41.99\ g\ NaF}{1\ mol\ NaF} = 5.25\ g\ NaF$$

Check. (0.5/4) = 0.125; 0.13 × 42 > 4 g NaF

(c) *Analyze*. Given: g HF Find: g Na_2SiO_3.

$$\text{Plan.}\ g\ HF \rightarrow mol\ HF \left(\begin{matrix} mol \\ ratio \end{matrix}\right) \rightarrow mol\ Na_2SiO_3 \rightarrow g\ Na_2SiO_3$$

The mole ratio is at the heart of every stoichiometry problem. Molar mass is used to change to and from grams. *Solve*.

$$0.800 \text{ g HF} \times \frac{1 \text{ mol HF}}{20.01 \text{ g HF}} \times \frac{1 \text{ mol Na}_2\text{SiO}_3}{8 \text{ mol HF}} \times \frac{122.1 \text{ g Na}_2\text{SiO}_3}{1 \text{ mol Na}_2\text{SiO}_3} = 0.610 \text{ g Na}_2\text{SiO}_3$$

Check. 0.8 (120/160) < 0.75 mol

3.59 (a) $Al(OH)_3(s) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)$

(b) *Analyze*. Given mass of one reactant, find stoichiometric mass of other reactant and products.

Plan. Follow the logic in Sample Exercise 3.16. Calculate mol $Al(OH)_3$ in 0.500 g $Al(OH_3)_3$ separately, since it will be used several times.

Solve. $$0.500 \text{ g Al(OH)}_3 \times \frac{1 \text{ mol Al(OH)}_3}{78.00 \text{ g Al(OH)}_3} = 6.410 \times 10^{-3} = 6.41 \times 10^{-3} \text{ mol Al(OH)}_3$$

$$6.410 \times 10^{-3} \text{ mol Al(OH)}_3 \times \frac{3 \text{ mol HCl}}{1 \text{ mol Al(OH)}_3} \times \frac{36.46 \text{ g HCl}}{1 \text{ mol HCl}} = 0.7012 = 0.701 \text{ g HCl}$$

(c) $$6.410 \times 10^{-3} \text{ mol Al(OH)}_3 \times \frac{1 \text{ mol HCl}}{1 \text{ mol Al(OH)}_3} \times \frac{133.34 \text{ g AlCl}_3}{1 \text{ mol AlCl}_3} = 0.8547$$

$$= 0.855 \text{ g AlCl}_3$$

$$6.410 \times 10^{-3} \text{ mol Al(OH)}_3 \times \frac{3 \text{ mol H}_2\text{O}}{1 \text{ mol Al(OH)}_3} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 0.3465 = 0.347 \text{ g H}_2\text{O}$$

(d) Conservation of mass: mass of products = mass of reactants

reactants: $Al(OH)_3$ + HCl, 0.500 g + 0.701 g = 1.201 g

products: $AlCl_3$ + H_2O, 0.855 g + 0.347 g = 1.202 g

The 0.001 g difference is due to rounding (0.8547 + 0.3465 = 1.2012). This is an excellent *check* of results.

3.61 (a) $Al_2S_3(s) + 6H_2O(l) \rightarrow 2Al(OH)_3(s) + 3H_2S(g)$

(b) *Plan*. g A → mol A → mol B → g B. See Solution 3.57 (c). *Solve*.

$$14.2 \text{ g Al}_2\text{S}_3 \times \frac{1 \text{ mol Al}_2\text{S}_3}{150.2 \text{ g Al}_2\text{S}_3} \times \frac{2 \text{ mol Al(OH)}_3}{1 \text{ mol Al}_2\text{S}_3} \times \frac{78.00 \text{ g Al(OH)}_3}{1 \text{ mol Al(OH)}_3} = 14.7 \text{ g Al(OH)}_3$$

Check. $$14\left(\frac{2 \times 78}{150}\right) \approx 14(1) \approx 14 \text{ g Al(OH)}_3$$

3.63 (a) *Analyze*. Given: mol NaN_3. Find: mol N_2.

Plan. Use mole ratio from balanced equation. *Solve*.

$$1.50 \text{ mol NaN}_3 \times \frac{3 \text{ mol N}_2}{2 \text{ mol NaN}_3} = 2.25 \text{ mol N}_2$$

Check. The resulting mol N_2 should be greater than mol NaN_3, (the N_2:NaN_3 ratio is > 1), and it is.

(b) *Analyze*. Given: g N_2 Find: g NaN_3.

Plan. Use molar masses to get from and to grams, mol ratio to relate moles of the two substances. *Solve*.

$$10.0 \text{ g } N_2 \times \frac{1 \text{ mol } N_2}{28.01 \text{ g } N_2} \times \frac{2 \text{ mol } NaN_3}{3 \text{ mol } N_2} \times \frac{65.01 \text{ g } NaN_3}{1 \text{ mol } NaN_3} = 15.5 \text{ g } NaN_3$$

Check. Mass relations are less intuitive than mole relations. Estimating the ratio of molar masses is sometimes useful. In this case, 65 g NaN_3/28 g N_2 ≈ 2.25 Then, (10 × 2/3 × 2.25) ≈ 14 g NaN_3. The calculated result looks reasonable.

(c) *Analyze*. Given: vol N_2 in ft^3, density N_2 in g/L. Find: g NaN_3.

Plan. First determine how many g N_2 are in 10.0 ft^3, using the density of N_2. Then proceed as in part (b).

Solve.

$$\frac{1.25 \text{ g}}{1 \text{ L}} \times \frac{1 \text{ L}}{1000 \text{ cm}^3} \times \frac{(2.54)^3 \text{ cm}^3}{1 \text{ in}^3} \times \frac{(12)^3 \text{ in}^3}{1 \text{ ft}^3} \times 10.0 \text{ ft}^3 = 354.0 = 354 \text{ g } N_2$$

$$354.0 \text{ g } N_2 \times \frac{1 \text{ mol } N_2}{28.01 \text{ g } N_2} \times \frac{2 \text{ mol } NaN_3}{3 \text{ mol } N_2} \times \frac{65.01 \text{ g } NaN_3}{1 \text{ mol } NaN_3} = 548 \text{ g } NaN_3$$

Check. 1 ft^3 ~ 28 L; 10 ft^3 ~ 280 L; 280 L × 1.25 ~ 350 g N_2

Using the ratio of molar masses from part (b), (350 × 2/3 × 2.25) ≈ 525 g NaN_3

3.65 (a) *Analyze*. Given: dimensions of Al foil. Find: mol Al.

$$\textit{Plan.}\quad \text{Dimensions} \rightarrow \text{vol} \xrightarrow{\text{density}} \text{mass} \xrightarrow[\text{mass}]{\text{molar}} \text{mol Al}$$

$$\textit{Solve.}\ 1.00 \text{ cm} \times 1.00 \text{ cm} \times 0.550 \text{ mm} \times \frac{1 \text{ cm}}{10 \text{ mm}} = 0.0550 \text{ cm}^3 \text{ Al}$$

$$0.0550 \text{ cm}^3 \text{ Al} \times \frac{2.699 \text{ g Al}}{1 \text{ cm}^3} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = 5.502 \times 10^{-3} = 5.50 \times 10^{-3} \text{ mol Al}$$

Check. 2.699/26.98 ≈ 0.1; (0.055 cm^3 × 0.1) = 5.5 × 10^{-3} mol Al

(b) *Plan*. Write the balanced equation to get a mole ratio; change mol Al → mol $AlBr_3$ → g $AlBr_3$.

Solve. $2Al(s) + 3Br_2(l) \rightarrow 2AlBr_3(s)$

$$5.502 \times 10^{-3} \text{ mol Al} \times \frac{2 \text{ mol } AlBr_3}{2 \text{ mol Al}} \times \frac{266.69 \text{ g } AlBr_3}{1 \text{ mol } AlBr_3} = 1.467 = 1.47 \text{ g } AlBr_3$$

Check. (0.006 × 1 × 270) ≈ 1.6 g $AlBr_3$

Limiting Reactants; Theoretical Yields

3.67 (a) The *limiting reactant* determines the maximum number of product moles resulting from a chemical reaction; any other reactant is an *excess reactant*.

(b) The limiting reactant regulates the amount of products, because it is completely used up during the reaction; no more product can be made when one of the reactants is unavailable.

(c) Combining ratios are molecule and mole ratios. Since different molecules have different masses, equal masses of different reactants will not have equal numbers of molecules. By comparing initial moles, we compare numbers of available reactant molecules, the fundamental combining units in a chemical reaction.

3.69 (a) Each bicycle needs 2 wheels, 1 frame, and 1 set of handlebars. A total of 4815 wheels corresponds to 2407.5 pairs of wheels. This is more than the number of frames or handlebars. The 2255 handlebars determine that 2255 bicycles can be produced.

(b) 2305 frames – 2255 bicycles = 50 frames left over

2407.5 pairs of wheels – 2255 bicycles = 152.5 pairs of wheels left over 2(152.5) = 305 wheels left over

(c) The handlebars are the "limiting reactant" in that they determine the number of bicycles that can be produced.

3.71 *Analyze*. Given: 1.85 mol NaOH, 1.00 mol CO_2. Find: mol Na_2CO_3.

Plan. Amounts of more than one reactant are given, so we must determine which reactant regulates (limits) product. Then apply the appropriate mole ratio from the balanced equation.

Solve. The mole ratio is 2NaOH:1CO_2, so 1.00 mol CO_2 requires 2.00 mol NaOH for complete reaction. Less than 2.00 mol NaOH are present, so NaOH is the limiting reactant.

$$1.85 \text{ mol NaOH} \times \frac{1 \text{ mol } Na_2CO_3}{2 \text{ mol NaOH}} = 0.925 \text{ mol } Na_2CO_3 \text{ can be produced}$$

The Na_2CO_3:CO_2 ratio is 1:1, so 0.925 mol Na_2CO_3 produced requires 0.925 mol CO_2 consumed. (Alternately, 1.85 mol NaOH × 1 mol CO_2/2 mol NaOH = 0.925 mol CO_2 reacted). 1.00 mol CO_2 initial – 0.925 mol CO_2 reacted = 0.075 mol CO_2 remain.

Check.	2NaOH(s)	+	CO_2(g)	→	Na_2CO_3(s)	+	H_2O(l)
initial	1.85 mol		1.00 mol		0 mol		
change (reaction)	–1.85 mol		–0.925 mol		+0.925 mol		
final	0 mol		0.075 mol		0.925 mol		

Note that the "change" line (but not necessarily the "final" line) reflects the mole ratios from the balanced equation.

3.73 $3NaHCO_3(aq) + H_3C_6H_5O_7(aq) \rightarrow 3CO_2(g) + 3H_2O(l) + Na_3C_6H_5O_7(aq)$

(a) *Analyze/Plan*. Abbreviate citric acid as H_3Cit. Follow the approach in Sample Exercise 3.19. *Solve*.

$$1.00 \text{ g } NaHCO_3 \times \frac{1 \text{ mol } NaHCO_3}{84.01 \text{ g } NaHCO_3} = 1.190 \times 10^{-2} = 1.19 \times 10^{-2} \text{ mol } NaHCO_3$$

$$1.00 \text{ g } H_3C_6H_5O_7 \times \frac{1 \text{ mol } H_3Cit}{192.1 \text{ g } H_3Cit} = 5.206 \times 10^{-3} = 5.21 \times 10^{-3} \text{ mol } H_3Cit$$

But $NaHCO_3$ and H_3Cit react in a 3:1 ratio, so 5.21×10^{-3} mol H_3Cit require $3(5.21 \times 10^{-3}) = 1.56 \times 10^{-2}$ mol $NaHCO_3$. We have only 1.19×10^{-2} mol $NaHCO_3$, so $NaHCO_3$ is the limiting reactant.

(b) $$1.190 \times 10^{-2} \text{ mol } NaHCO_3 \times \frac{3 \text{ mol } CO_2}{3 \text{ mol } NaHCO_3} \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 0.524 \text{ g } CO_2$$

(c) $$1.190 \times 10^{-2} \text{ mol } NaHCO_3 \times \frac{1 \text{ mol } H_3Cit}{3 \text{ mol } NaHCO_3} = 3.968 \times 10^{-3}$$
$$= 3.97 \times 10^{-3} \text{ mol } H_3Cit \text{ react}$$

5.206×10^{-3} mol H_3Cit – 3.968×10^{-3} mol react = 1.238×10^{-3}

= 1.24×10^{-3} mol H_3Cit remain

$$1.238 \times 10^{-3} \text{ mol } H_3Cit \times \frac{192.1 \text{ g } H_3Cit}{\text{mol } H_3Cit} = 0.238 \text{ g } H_3Cit \text{ remain}$$

3.75 *Analyze*. Given: initial g Na_2CO_3, g $AgNO_3$. Find: final g Na_2CO_3, $AgNO_3$, Ag_2CO_3, $NaNO_3$

Plan. Write balanced equation; determine limiting reactant; calculate amounts of excess reactant remaining and products, based on limiting reactant.

Solve. $2AgNO_3(aq) + Na_2CO_3(aq) \rightarrow Ag_2CO_3(s) + 2NaNO_3(aq)$

$$3.50 \text{ g } Na_2CO_3 \times \frac{1 \text{ mol } Na_2CO_3}{106.0 \text{ g } Na_2CO_3} = 0.03302 = 0.0330 \text{ mol } Na_2CO_3$$

$$5.00 \text{ g } AgNO_3 \times \frac{1 \text{ mol } AgNO_3}{169.9 \text{ g } AgNO_3} = 0.02943 = 0.0294 \text{ mol } AgNO_3$$

$$0.02943 \text{ mol } AgNO_3 \times \frac{1 \text{ mol } Na_2CO_3}{2 \text{ mol } AgNO_3} = 0.01471 = 0.0147 \text{ mol } Na_2CO_3 \text{ required}$$

$AgNO_3$ is the limiting reactant and Na_2CO_3 is present in excess.

	$2AgNO_3(aq)$	+	$Na_2CO_3(aq)$	$\rightarrow$	$Ag_2CO_3(s)$	+	$2NaNO_3(aq)$
initial	0.0294 mol		0.0330 mol		0 mol		0 mol
reaction	–0.0294 mol		–0.0147 mol		+0.0147 mol		+0.0294 mol
final	0 mol		0.0183 mol		0.0147 mol		0.0294 mol

0.01830 mol Na_2CO_3 × 106.0 g/mol = 1.940 = 1.94 g Na_2CO_3

0.01471 mol Ag_2CO_3 × 275.8 g/mol = 4.057 = 4.06 g Ag_2CO_3

0.02943 mol $NaNO_3$ × 85.00 g/mol = 2.502 = 2.50 g $NaNO_3$

Check. The initial mass of reactants was 8.50 g, and the final mass of excess reactant and products is 13.50 g; mass is conserved.

3.77 *Analyze*. Given: amounts of two reactants. Find: theoretical yield.

Plan. Determine the limiting reactant and the maximum amount of product it could produce. Then calculate % yield. *Solve*.

(a) $$30.0\text{ g } C_6H_6 \times \frac{1\text{ mol } C_6H_6}{78.11\text{ g } C_6H_6} = 0.3841 = 0.384\text{ mol } C_6H_6$$

$$65.0\text{ g } Br_2 \times \frac{1\text{ mol } Br_2}{159.8\text{ g } Br_2} = 0.4068 = 0.407\text{ mol } Br_2$$

Since C_6H_6 and Br_2 react in a 1:1 mole ratio, C_6H_6 is the limiting reactant and determines the theoretical yield.

$$0.3841\text{ mol } C_6H_6 \times \frac{1\text{ mol } C_6H_5Br}{1\text{ mol } C_6H_6} \times \frac{157.0\text{ g } C_6H_5Br}{1\text{ mol } C_6H_5Br} = 60.30 = 60.3\text{ g } C_6H_5Br$$

Check. 30/78 ~ 3/8 mol C_6H_6. 65/160 ~ 3/8 mol Br_2. Since moles of the two reactants are similar, a precise calculation is needed to determine the limiting reactant. 3/8 × 160 ≈ 60 g product

(b) $$\%\text{ yield} = \frac{42.3\text{ g } C_6H_5Br\text{ actual}}{60.3\text{ g } C_6H_5Br\text{ theoretical}} \times 100 = 70.149 = 70.1\%$$

3.79 *Analyze*. Given: g of two reactants, % yield. Find: g S_8.

Plan. Determine limiting reactant and theoretical yield. Use definition of % yield to calculate actual yield. *Solve*.

$$30.0\text{ g } H_2S \times \frac{1\text{ mol } H_2S}{34.08\text{ g } H_2S} = 0.8803 = 0.880\text{ mol } H_2S$$

$$50.0\text{ g } O_2 \times \frac{1\text{ mol } O_2}{32.00\text{ g } H_2S} = 1.5625 = 1.56\text{ mol } O_2$$

$$0.8803\text{ mol } H_2S \times \frac{4\text{ mol } O_2}{8\text{ mol } H_2S} = 0.4401 = 0.440\text{ mol } O_2\text{ required}$$

Since there is more than enough O_2 to react exactly with 0.880 mol H_2S, O_2 is present in excess and H_2S is the limiting reactant.

$$0.8803\text{ mol } H_2S \times \frac{1\text{ mol } S_8}{8\text{ mol } H_2S} \times \frac{256.56\text{ g } S_8}{1\text{ mol } S_8} = 28.231 = 28.2\text{ g } S_8\text{ theoretical yield}$$

Check. 30/34 ≈ 1 mol H_2S; 50/32 ≈ 1.5 mol O_2. Twice as many mol H_2S as mol O_2 are required, so H_2S limits. 1 × (260/8) ≈ 30 g S_8 theoretical.

$$\%\text{ yield} = \frac{\text{actual}}{\text{theoretical}} \times 100;\quad \frac{\%\text{ yield} \times \text{theoretical}}{100} = \text{actual yield}$$

$$\frac{98\%}{100} \times 28.231\text{ g } S_8 = 27.666 = 28\text{ g } S_8\text{ actual}$$

Additional Exercises

3.81 (a) $CH_3COOH = C_2H_4O_2$. At room temperature and pressure, pure acetic acid is a liquid. $C_2H_4O_2(l) + 2\ O_2(g) \rightarrow 2\ CO_2(g) + 2\ H_2O(l)$

(b) $Ca(OH)_2(s) \rightarrow CaO(s) + H_2O(g)$

(c) $Ni(s) + Cl_2(g) \rightarrow NiCl_2(s)$

3.85 (a) *Analyze*. Given: diameter of Si sphere (dot), density of Si. Find: mass of dot.

Plan. Calculate volume of sphere in cm^3, use density to calculate mass of the sphere (dot).

Solve. $V = 4/3\pi\, r^3$; $r = d/2$

$$\text{radius of dot} = \frac{4\ \text{nm}}{2} \times \frac{1 \times 10^{-9}\ \text{m}}{1\ \text{nm}} \times \frac{1\ \text{cm}}{1 \times 10^{-9}\ \text{m}} = 2 \times 10^{-7}\ \text{cm}$$

$$\text{volume of dot} = (4/3) \times \pi \times (2 \times 10^{-7})^3 = 3.35 \times 10^{-20} = 3 \times 10^{-20}\ \text{cm}^3$$

$$3.35 \times 10^{-20}\ \text{cm}^3 \times \frac{2.33\ \text{g Si}}{\text{cm}^3} = 7.707 \times 10^{-20} = 8 \times 10^{-20}\ \text{g Si in dot}$$

(b) *Plan*. Change g Si to mol Si using molar mass, then mol Si to atoms Si using Avogadro's number. *Solve*.

$$7.707 \times 10^{-20}\ \text{g Si} \times \frac{1\ \text{mol Si}}{28.0855\ \text{g Si}} \times \frac{6.022 \times 10^{23}\ \text{Si atoms}}{\text{mol Si}} = 1.653 \times 10^3$$

$$= 2 \times 10^3\ \text{Si atoms}$$

(c) *Plan*. A 4 nm quantum dot of Ge also has a volume of $3 \times 10^{-20}\ cm^3$. Use density of Ge and Avogadro's number to calculate the number of Ge atoms in a 4 nm spherical quantum dot.

$$3.35 \times 10^{-20}\ \text{cm}^3 \times \frac{5.325\ \text{g Ge}}{\text{cm}^3} \times \frac{1\ \text{mol Ge}}{72.64\ \text{g Ge}} \times \frac{6.022 \times 10^{23}\ \text{Ge atoms}}{\text{mol Ge}}$$

$$= 1.479 \times 10^3 = 1 \times 10^3\ \text{Ge atoms}$$

Strictly speaking, the result has 1 sig fig (from 4 nm). A more meaningful comparison might be 1700 Si atoms vs. 1500 Ge atoms. Although Ge has greater molar mass, it is also more than twice as dense as Si, so the numbers of atoms in the Si and Ge dots are similar.

3.89 *Plan*. Because different sample sizes were used to analyze the different elements, calculate mass % of each element in the sample.

i. Calculate mass % C from g CO_2.

ii. Calculate mass % Cl from AgCl.

iii. Get mass % H by subtraction.

iv. Calculate mole ratios and the empirical formulas.

Solve.

i. $3.52\text{ g CO}_2 \times \frac{1\text{ mol CO}_2}{44.01\text{ g CO}_2} \times \frac{1\text{ mol C}}{1\text{ mol CO}_2} \times \frac{12.01\text{ g C}}{1\text{ mol C}} = 0.9606 = 0.961\text{ g C}$

$\frac{0.9606\text{ g C}}{1.50\text{ g sample}} \times 100 = 64.04 = 64.0\%\text{ C}$

ii. $1.27\text{ g AgCl} \times \frac{1\text{ mol AgCl}}{143.3\text{ g AgCl}} \times \frac{1\text{ mol Cl}}{1\text{ mol AgCl}} \times \frac{35.45\text{ g Cl}}{1\text{ mol Cl}} = 0.3142 = 0.314\text{ g Cl}$

$\frac{0.3142\text{ g Cl}}{1.00\text{ g sample}} \times 100 = 31.42 = 31.4\%\text{ Cl}$

iii. % H = 100.0 – (64.04% C + 31.42% Cl) = 4.54 = 4.5% H

iv. Assume 100 g sample.

$64.04\text{ g C} \times \frac{1\text{ mol C}}{12.01\text{ g C}} = 5.33\text{ mol C}; \quad 5.33/0.886 = 6.02$

$31.42\text{ g Cl} \times \frac{1\text{ mol Cl}}{35.45\text{ g Cl}} = 0.886\text{ mol Cl}; \quad 0.886/0.886 = 1.00$

$4.54\text{ g H} \times \frac{1\text{ mol H}}{1.008\text{ g H}} = 4.50\text{ mol H}; \quad 4.50/0.886 = 5.08$

The empirical formula is probably C_6H_5Cl.

The subscript for H, 5.08, is relatively far from 5.00, but C_6H_5Cl makes chemical sense. More significant figures in the mass data are required for a more accurate mole ratio.

3.93 $O_3(g) + 2NaI(aq) + H_2O(l) \rightarrow O_2(g) + I_2(s) + 2NaOH(aq)$

(a) $5.95 \times 10^{-6}\text{ mol O}_3 \times \frac{2\text{ mol NaI}}{1\text{ mol O}_3} = 1.19 \times 10^{-5}\text{ mol NaI}$

(b) $1.3\text{ mg O}_3 \times \frac{1 \times 10^{-3}\text{ g}}{1\text{ mg}} \times \frac{1\text{ mol O}_3}{48.00\text{ g O}_3} \times \frac{2\text{ mol NaI}}{1\text{ mol O}_3} \times \frac{149.9\text{ g NaI}}{1\text{ mol NaI}}$

$= 8.120 \times 10^{-3} = 8.1 \times 10^{-3}\text{ g NaI} = 8.1\text{ mg NaI}$

3.97 $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

Determine the moles of N_2 and H_2 required to form the 3.0 moles of NH_3 present after the reaction has stopped.

$3.0\text{ mol NH}_3 \times \frac{3\text{ mol H}_2}{2\text{ mol NH}_3} = 4.5\text{ mol H}_2\text{ reacted}$

$3.0\text{ mol NH}_3 \times \frac{1\text{ mol N}_2}{2\text{ mol NH}_3} = 1.5\text{ mol N}_2\text{ reacted}$

mol H_2 initial = 3.0 mol H_2 remain + 4.5 mol H_2 reacted = 7.5 mol H_2

mol N_2 initial = 3.0 mol N_2 remain + 1.5 mol N_2 reacted = 4.5 mol N_2

In tabular form: $N_2(g) \quad + \quad 3H_2(g) \quad \rightarrow \quad 2NH_3(g)$

initial	4.5 mol	7.5 mol	0 mol
reaction	–1.5 mol	–4.5 mol	+3.0 mol
final	3.0 mol	3.0 mol	3.0 mol

(Tables like this will be extremely useful for solving chemical equilibrium problems in Chapter 15.)

Integrative Exercises

3.101 *Plan.* $\text{Volume cube} \xrightarrow{\text{density}} \text{mass } CaCO_3 \rightarrow \text{moles } CaCO_3 \rightarrow \text{moles O} \rightarrow \text{O atoms}$

Solve $(2.005)^3 \text{ in}^3 \times \frac{(2.54)^3 \text{ cm}^3}{1 \text{ in}^3} \times \frac{2.71 \text{ g } CaCO_3}{1 \text{ cm}^3} \times \frac{1 \text{ mol } CaCO_3}{100.1 \text{ g } CaCO_3} \times \frac{3 \text{ mol O}}{1 \text{ mol } CaCO_3}$

$\times \frac{6.022 \times 10^{23} \text{ O atoms}}{1 \text{ mol O}} = 6.46 \times 10^{24} \text{ O atoms}$

3.105 (a) $S(s) + O_2(g) \rightarrow SO_2(g)$; $SO_2(g) + CaO(s) \rightarrow CaSO_3(s)$

(b) If the coal contains 2.5% S, then 1 g coal contains 0.025 g S.

$$\frac{2000 \text{ tons coal}}{\text{day}} \times \frac{2000 \text{ lb}}{1 \text{ ton}} \times \frac{1 \text{ kg}}{2.205 \text{ lb}} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{0.025 \text{ g S}}{1 \text{ g coal}} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}}$$

$$\times \frac{1 \text{ mol } SO_2}{1 \text{ mol S}} \times \frac{1 \text{ mol CaO}}{1 \text{ mol } SO_2} \times \frac{56.08 \text{ g CaO}}{1 \text{ mol CaO}} \times \frac{1 \text{ kg CaO}}{1000 \text{ g CaO}} =$$

$$79{,}485 = 7.9 \times 10^4 \text{ kg CaO or } 7.9 \times 10^7 \text{ g CaO}$$

(c) 1 mol CaO = 1 mol $CaSO_3$

$$7.9485 \times 10^7 \text{ g CaO} \times \frac{1 \text{ mol CaO}}{56.08 \text{ g CaO}} \times \frac{1 \text{ mol } CaSO_3}{1 \text{ mol CaO}} \times \frac{120.14 \text{ g } CaSO_3}{1 \text{ mol } CaSO_3}$$

$$= 1.7028 \times 10^8 = 1.7 \times 10^8 \text{ g } CaSO_3$$

This corresponds to about 190 tons of $CaSO_3$ per day as a waste product.