3. CALCULATIONS WITH CHEMICAL FORMULAS AND EQUATIONS

■ Solutions to Exercises

Note on significant figures: If the final answer to a solution needs to be rounded off, it is given first with one nonsignificant figure, and the last significant figure is underlined. The final answer is then rounded to the correct number of significant figures. In multiple-step problems, intermediate answers are given with at least one nonsignificant figure; however, only the final answer has been rounded off.

3.1 a. NO_2

1 x AW of N			=	14.0067 amu
2 x AW of O	=	2 x 15.9994	=	31.9988 amu
		MW of NO_2	=	46.0055 = 46.0 amu (3 s.f.)

b. $C_6H_{12}O_6$

6 x AW of C	=	6 x 12.011	=	72.066 amu
12 x AW of H	=	12 x 1.0079	=	12.0948 amu
6 x AW of O	=	6 x 15.9994	=	95.9964 amu
		MW of $C_6H_{12}O_6$	=	180.1572 amu = 180. amu (3 s.f.)

c. NaOH

1 x AW of Na			=	22.98977 amu
1 x AW of O			=	15.9994 amu
1 x AW of H			=	1.0079 amu
		MW of NaOH	=	39.9971 amu = 40.0 amu (3 s.f.)

d. $Mg(OH)_2$

1 x AW of Mg			=	24.305 amu
2 x AW of O	=	2 x 15.9994	=	31.9988 amu
2 x AW of H	=	2 x 1.0079	=	2.0158 amu
		MW of $Mg(OH)_2$	=	58.3196 amu = 58.3 amu (3 s.f.)

3.2 a. The molecular model represents a molecule made up of one S and three O. The chemical formula is SO_3. Using the same approach as Example 3.1 in the text, calculating the formula weight yields 80.07 amu.

b. The molecular model represents one S, four O, and two H. The chemical formula is then H_2SO_4. The formula weight is 98.09 amu.

3.3 a. The atomic weight of Ca = 40.08 amu; thus, the molar mass = 40.08 g/mol, and one mol of Ca = 6.022×10^{23} Ca atoms.

$$\text{Mass of one Ca} = \frac{40.08\text{ g}}{1\text{ mol Ca}} = \frac{1\text{ mol}}{6.022 \times 10^{23}\text{ atoms}} = 6.65\underline{5}6 \times 10^{-23}$$

$$= 6.656 \times 10^{-23}\text{ g/atom}$$

b. The molecular weight of C_2H_5OH, or C_2H_6O, = (2 x 12.01) + (6 x 1.008) + 16.00 = 46.0$\underline{6}$8. Its molar mass = 46.07 g/mol, and one mol = 6.022×10^{23} molecules of C_2H_6O.

$$\text{Mass of one } C_2H_6O = \frac{46.07\text{ g}}{1\text{ mol } C_2H_6O} = \frac{1\text{ mol}}{6.022 \times 10^{23}\text{ atoms}}$$

$$= 7.65\underline{0}3 \times 10^{-23} = 7.650 \times 10^{-23}\text{ g/molecule}$$

3.4 The molar mass of H_2O_2 is 34.02 g/mol. Therefore,

$$0.909\text{ mol } H_2O_2 \times \frac{34.02\text{ g } H_2O_2}{1\text{ mol } H_2O_2} = 30.\underline{9}2 = 30.9\text{ g } H_2O_2$$

3.5 The molar mass of HNO_3 is 63.01 g/mol. Therefore,

$$28.5\text{ g } HNO_3 \times \frac{1\text{ mol } HNO_3}{63.01\text{ g } HNO_3} = 0.45\underline{2}3 = 0.452\text{ mol } HNO_3$$

3.6 Convert the mass of HCN from milligrams to grams. Then, convert grams of HCN to moles of HCN. Finally, convert moles of HCN to the number of HCN molecules.

$$56\text{ mg HCN} \times \frac{1\text{ g}}{1000\text{ mg}} \times \frac{1\text{ mol HCN}}{27.02\text{ g HCN}} \times \frac{6.022 \times 10^{23}\text{ HCN molecules}}{1\text{ mol HCN}}$$

$$= 1.\underline{2}48 \times 10^{21} = 1.2 \times 10^{21}\text{ HCN molecules}$$

3.7 The molecular weight of NH_4NO_3 = 80.05; thus, its molar mass = 80.05 g/mol. Hence

$$\text{Percent N} = \frac{28.02\text{ g}}{80.05\text{ g}} \times 100\% = 35.\underline{0}0 = 35.0\%$$

$$\text{Percent H} = \frac{4.032\text{ g}}{80.05\text{ g}} \times 100\% = 5.0\underline{3}6 = 5.04\%$$

$$\text{Percent O} = \frac{48.00\text{ g}}{80.05\text{ g}} \times 100\% = 59.\underline{9}6 = 60.0\%$$

3.8 From the previous exercise, NH_4NO_3 is 35.0 percent N (fraction N = 0.350), so the mass of N in 48.5 g of NH_4NO_3 is

$$48.5\text{ g } NH_4NO_3 \times (0.350\text{ g N}/1\text{ g } NH_4NO_3) = 16.\underline{9}75 = 17.0\text{ g N}$$

3.9 First, convert the mass of CO_2 to moles of CO_2. Next, convert this to moles of C (one mol of CO_2 is equivalent to one mol C). Finally, convert to mass of carbon, changing mg to g first:

$$5.80 \times 10^{-3}\text{ g } CO_2 \times \frac{1\text{ mol } CO_2}{44.01\text{ g } CO_2} \times \frac{1\text{ mol C}}{1\text{ mol } CO_2} \times \frac{12.01\text{ g C}}{1\text{ mol C}}$$
$$= 1\ 5\underline{8}3 \times 10^{-3}\text{ g C}$$

Do the same series of calculations for water, noting that one mol H_2O contains two mol H.

$$1.58 \times 10^{-3}\text{ g } H_2O \times \frac{1\text{ mol } H_2O}{18.02\text{ g } H_2O} \times \frac{2\text{ mol H}}{1\text{ mol } H_2O} \times \frac{1.008\text{ g H}}{1\text{ mol H}}$$
$$= 1.7\underline{6}7 \times 10^{-4}\text{ g H}$$

The mass percentages of C and H can be calculated using the masses from the previous calculations:

$$\text{Percent C} = \frac{1.583\text{ mg}}{3.87\text{ mg}} \times 100\% = 40.\underline{9}0 = 40.9\%\text{ C}$$

$$\text{Percent H} = \frac{0.1767\text{ mg}}{3.87\text{ mg}} \times 100\% = 4.5\underline{6}58 = 4.57\%\text{ H}$$

The mass percentage of O can be determined by subtracting the sum of the above percentages from 100 percent:

$$\text{Percent O} = 100.000\% - (40.90 + 4.5658) = 54.\underline{5}342 = 54.5\%\text{ O}$$

3.10 Convert the masses to moles that are proportional to the subscripts in the empirical formula:

$$33.4 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = 1.0\underline{4}14 \text{ mol S}$$

$$(83.5 - 33.4) \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.1\underline{3}2 \text{ mol O}$$

Next, obtain the smallest integers from the moles by dividing each by the smallest number of moles:

$$\text{For O: } \frac{3.1312}{1.0414} = 3.01 \qquad \text{For S: } \frac{1.0414}{1.0414} = 1.00$$

The empirical formula is SO_3.

3.11 For a 100.0-g sample of benzoic acid, 68.8 g are C, 5.0 g are H, and 26.2 g are O. Using the molar masses, convert these masses to moles:

$$68.8 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 5.7\underline{2}9 \text{ mol C}$$

$$5.0 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.\underline{9}6 \text{ mol H}$$

$$26.2 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.6\underline{3}8 \text{ mol O}$$

These numbers are in the same ratio as the subscripts in the empirical formula. They must be changed to integers. First, divide each one by the smallest mol number:

$$\text{For C}: \frac{5.729}{1.638} = 3.497 \qquad \text{For H}: \frac{4.96}{1.638} = 3.03$$

$$\text{For O}: \frac{1.638}{1.638} = 1.000$$

Rounding off, we obtain $C_{3.5}H_{3.0}O_{1.0}$. Multiplying the numbers by two gives whole numbers for an empirical formula of $C_7H_6O_2$.

3.12 For a 100.0-g sample of acetaldehyde, 54.5 g are C, 9.2 g are H, and 36.3 g are O. Using the molar masses, convert these masses to moles:

$$54.5 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.5\underline{3}7 \text{ mol C}$$

$$9.2 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 9.\underline{1}2 \text{ mol H}$$

$$36.3 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.2\underline{6}8 \text{ mol O}$$

These numbers are in the same ratio as the subscripts in the empirical formula. They must be changed to integers. First, divide each one by the smallest mol number:

$$\text{For C}: \frac{4.537}{2.268} = 2.000 \qquad \text{For H}: \frac{9.12}{2.268} = 4.02$$

$$\text{For O}: \frac{2.268}{2.268} = 1.000$$

Rounding off, we obtain C_2H_4O, the empirical formula, which is also the molecular formula.

3.13

H_2	+	Cl_2	$\rightarrow$	2HCl
1 molec. (mol) H_2	+	1 molec. (mol) Cl_2	$\rightarrow$	2 molec. (mol) HCl (molec., mole interp.)
2.016 g H_2	+	70.9 g Cl_2	$\rightarrow$	2 x 36.5 g HCl (mass interp.)

3.14 Equation: $Na + H_2O \rightarrow 1/2H_2 + NaOH$, or $2Na + 2H_2O \rightarrow H_2 + 2NaOH$. From this equation, one mol of Na corresponds to one-half mol of H_2, or two mol of Na corresponds to one mol of H_2. Therefore,

$$7.81 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.016 \text{ g } H_2} \times \frac{2 \text{ mol Na}}{1 \text{ mol } H_2} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 17\underline{8}.1 = 178 \text{ g Na}$$

3.15 Balanced equation: $2ZnS + 3O_2 \rightarrow 2ZnO + 2SO_2$

Convert grams of ZnS to moles of ZnS. Then, determine the relationship between ZnS and O_2 ($2ZnS$ is equivalent to $3O_2$). Finally, convert to mass O_2.

$$5.00 \times 10^3 \text{ g ZnS} \times \frac{1 \text{ mol ZnS}}{97.46 \text{ g ZnS}} \times \frac{3 \text{ mol } O_2}{2 \text{ mol ZnS}} \times \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} \times \frac{1 \text{ kg}}{1000 \text{ g}}$$

$$= 2.4\underline{6}3 = 2.46 \text{ kg } O_2$$

3.16 Balanced equation: $2HgO \rightarrow 2Hg + O_2$

Convert the mass of O_2 to mol of O_2. Using the fact that one mol of O_2 is equivalent to two mol of Hg, determine the number of mol of Hg, and convert to mass of Hg.

$$6.47 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} \times \frac{2 \text{ mol Hg}}{1 \text{ mol } O_2} \times \frac{200.59 \text{ g Hg}}{1 \text{ mol Hg}} = 81.\underline{1}1$$

$$= 81.1 \text{ g Hg}$$

3.17 First, determine the limiting reactant by calculating the moles of $AlCl_3$ that would be obtained if Al and HCl were totally consumed:

$$0.15 \text{ mol Al} \times \frac{2 \text{ mol } AlCl_3}{2 \text{ mol Al}} = 0.1\underline{5}0 \text{ mol } AlCl_3$$

$$0.35 \text{ mol HCl} \times \frac{2 \text{ mol } AlCl_3}{6 \text{ mol HCl}} = 0.1\underline{1}66 \text{ mol } AlCl_3$$

Because the HCl produces the smaller amount of $AlCl_3$, the reaction will stop when HCl is totally consumed but before the Al is consumed. The limiting reactant is therefore HCl. The amount of $AlCl_3$ produced must be $0.1\underline{1}66$, or 0.12 mol.

3.18 First, determine the limiting reactant by calculating the moles of ZnS produced by totally consuming Zn and S_8:

$$7.36 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.39 \text{ g Zn}} \times \frac{8 \text{ mol ZnS}}{8 \text{ mol Zn}} = 0.11\underline{2}56 \text{ mol ZnS}$$

$$6.45 \text{ g } S_8 \times \frac{1 \text{ mol } S_8}{256.56 \text{ g } S_8} \times \frac{8 \text{ mol ZnS}}{1 \text{ mol } S_8} = 0.20\underline{1}1 \text{ mol ZnS}$$

The reaction will stop when Zn is totally consumed; S_8 is in excess, and not all of it is converted to ZnS. The limiting reactant is therefore Zn. Now convert the moles of ZnS obtained from the Zn to grams of ZnS:

$$0.11256 \text{ mol ZnS} \times \frac{97.46 \text{ g ZnS}}{1 \text{ mol ZnS}} = 10.\underline{9}7 = 11.0 \text{ g ZnS}$$

3.19 First, write the balanced equation:

$$CH_3OH + CO \rightarrow HC_2H_3O_2$$

Convert grams of each reactant to moles of acetic acid:

$$15.0\text{ g } CH_3OH \times \frac{1\text{ mol } CH_3OH}{32.04\text{ g } CH_3OH} \times \frac{1\text{ mol } HC_2H_3O_2}{1\text{ mol } CH_3OH} = 0.46\underline{8}1\text{ mol } HC_2H_3O_2$$

$$10.0\text{ g CO} \times \frac{1\text{ mol CO}}{28.01\text{ g CO}} \times \frac{1\text{ mol } HC_2H_3O_2}{1\text{ mol } CH_3OH} = 0.35\underline{7}0\text{ mol } HC_2H_3O_2$$

Thus, CO is the limiting reactant, and 0.03570 mol $HC_2H_3O_2$ is obtained. The mass of product is

$$0.3570\text{ mol } HC_2H_3O_2 \times \frac{60.05\text{ g } HC_2H_3O_2}{1\text{ mol } HC_2H_3O_2} = 21.\underline{4}4\text{ g } HC_2H_3O_2$$

The percentage yield is:

$$\frac{19.1\text{ g actual yield}}{21.44\text{ g theoretical yield}} \times 100\% = 89.\underline{0}8 = 89.1\%$$

$$\text{Mass percentage} = \frac{0.25359\text{ g } HC_2H_3O_2}{5.00\text{ g vinegar}} \times 100\% = 5.0\underline{7}1 = 5.07\%$$

■ Answers to Concept Checks

3.1 a. Each tricycle has one seat, so you have a total of 1.5 mol of seats.

b. Each tricycle has three tires, so you have 1.5 mol x 3 = 4.5 mol of tires.

c. Each $Mg(OH)_2$ has two OH^- ions, so there are 1.5 mol x 2 = 3.0 mol OH^- ions.

3.2 a. When conducting this type of experiment, you are assuming that all of the carbon and hydrogen show up in the CO_2 and H_2O, respectively. In this experiment, where all of the carbon and hydrogen do not show up, when you analyze the CO_2 for carbon and H_2O for hydrogen, you find the weights in the products are less than those in the carbon and hydrogen you started with.

b. Since you collected less carbon and hydrogen than were present in the original sample, the calculated mass percentage will be less than the expected (real) value. For example, say you have a 10.0 g sample that contains 7.5 g of carbon. You run the experiment on the 10.0 g sample and collect only 5.0 g of carbon. The calculated percent carbon based on your experimental results would be 50 percent instead of the correct amount of 75 percent.

3.3 a. $C_2H_8O_2$ is not an empirical formula because each of the subscripts can be divided by two to obtain a possible empirical formula of CH_4O. (The empirical formula is not the <u>smallest</u> integer ratio of subscripts.)

b. $C_{1.5}H_4$ is not a correct empirical formula because one of the subscripts is not an integer. Multiply each of the subscripts by two to obtain the possible empirical formula C_3H_8. (Since the subscript of carbon is the decimal number 1.5, the empirical formula is not the smallest <u>integer</u> ratio of subscripts.)

c. Yes, the empirical formula and the molecular formula can be the same, as is the case in this problem where the formula is written with the smallest integer subscripts.

3.4 a. Correct. Coefficients in balanced equations can represent amounts in atoms and molecules.

b. Incorrect. The coefficients in a balanced chemical equation do not represent amounts in grams. One gram of carbon and one gram of oxygen represent different molar amounts.

c. Incorrect. The coefficients in a balanced chemical equation do not represent amounts in grams.

d. Correct. You might initially think this is an incorrect representation; however, 12g of C, 32 g of O_2, and 44 g of CO_2 each represent one mole of the substance, so the relationship of the chemical equation is obeyed.

e. Correct. The coefficients in balanced equations can represent amounts in moles.

f. Incorrect. The amount of O_2 present is not enough to react completely with one mol of carbon. Only one-half of the carbon would react, and one-half mol of CO_2 would form.

g. Incorrect. In this representation, oxygen is being shown as individual atoms of O, not as molecules of O_2, so the drawings are not correctly depicting the chemical reaction.

h. Correct. The molecular models correctly depict a balanced chemical reaction since the same number of atoms of each element appears on both sides of the equation.

3.5 a. $X_2(g) + 2Y(g) \rightarrow 2XY(g)$

(continued)

b. Since the product consists of a combination of X and Y in a 1:1 ratio, it must consist of two atoms hooked together. If you count the total number of X atoms (split apart the X_2 molecules) and Y atoms present prior to the reaction, there are four X atoms and three Y atoms. From these starting quantities, you are limited to three XY molecules and left with an unreacted Y. Option #1 represents this situation and is therefore the correct answer.

c. Since X_2(g) was completely used up during the course of the reaction, it is the limiting reactant.

■ Answers to Review Questions

3.1 The molecular weight is the sum of the atomic weights of all the atoms in a molecule of the substance whereas the formula weight is the sum of the atomic weights of all the atoms in one formula unit of the compound, whether the compound is molecular or not. A given substance could have both a molecular weight and a formula weight if it existed as discrete molecules.

3.2 To obtain the formula weight of a substance, sum up the atomic weights of all atoms in the formula.

3.3 A mole of N_2 contains Avogadro's number (6.02×10^{23}) of N_2 molecules and $2 \times 6.02 \times 10^{23}$ of N atoms. One mole of $Fe_2(SO_4)_3$ contains three moles of SO_4^{2-} ions; and it contains twelve moles of O atoms.

3.4 A sample of the compound of known mass is burned, and CO_2 and H_2O are obtained as products. Next, you relate the masses of CO_2 and H_2O to the masses of carbon and hydrogen. Then, you calculate the mass percentages of C and H. You find the mass percentage of O by subtracting the mass percentages of C and H from 100.

3.5 The empirical formula is obtained from the percentage composition by assuming for the purposes of the calculation a sample of 100 g of the substance. Then, the mass of each element in the sample equals the numerical value of the percentage. Convert the masses of the elements to moles of the elements using the atomic mass of each element. Divide the moles of each by the smallest number to obtain the smallest ratio of each atom. If necessary, find a whole-number factor to multiply these results by to obtain integers for the subscripts in the empirical formula.

3.6 The empirical formula is the formula of a substance written with the smallest integer (whole number) subscripts. Each of the subscripts in the formula $C_6H_{12}O_2$ can be divided by two, so the empirical formula of the compound is C_3H_6O.

3.7 The number of empirical formula units in a compound, n, equals the molecular weight divided by the empirical formula weight.

$$n = \frac{34.0 \text{ amu}}{17.0 \text{ amu}} = 2.00$$

The molecular formula of hydrogen peroxide is therefore $(HO)_2$, or H_2O_2.

3.8 The coefficients in a chemical equation can be interpreted directly in terms of molecules or moles. For the mass interpretation, you will need the molar masses of CH_4, O_2, CO_2, and H_2O, which are 16.0, 32.0, 44.0, and 18.0 g/mol, respectively. A summary of the three interpretations is given below the balanced equation:

CH_4	+	$2O_2$	$\rightarrow$	CO_2	+	$2H_2O$
1 molecule	+	2 molecules	$\rightarrow$	1 molecule	+	2 molecules
1 mole	+	2 moles	$\rightarrow$	1 mole	+	2 moles
16.0 g	+	2 x 32.0 g	$\rightarrow$	44.0 g	+	2 x 18.0 g

3.9 A chemical equation yields the mole ratio of a reactant to a second reactant or product. Once the mass of a reactant is converted to moles, this can be multiplied by the appropriate mole ratio to give the moles of a second reactant or product. Multiplying this number of moles by the appropriate molar mass gives mass. Thus, the masses of two different substances are related by a chemical equation.

3.10 The limiting reactant is the reactant that is entirely consumed when the reaction is complete. Because the reaction stops when the limiting reactant is used up, the moles of product are always determined by the starting number of moles of the limiting reactant.

3.11 Two examples are given in the book. The first involves making cheese sandwiches. Each sandwich requires two slices of bread and one slice of cheese. The limiting reactant is the cheese because some bread is left unused. The second example is assembling automobiles. Each auto requires one steering wheel, four tires, and other components. The limiting reactant is the tires, since they will run out first.

3.12 Since the theoretical yield represents the maximum amount of product that can be obtained by a reaction from given amounts of reactants under any conditions, in an actual experiment you can never obtain more than this amount.

■ Answers to Conceptual Problems

3.13 a. $3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$

b. Since there is no H_2 present in the container, it was entirely consumed during the reaction, which makes it the limiting reactant.

c. According to the chemical reaction, three molecules of H_2 are required for every molecule of N_2. Since there are two moles of unreacted N_2, you would need six additional moles of H_2 to complete the reaction.

3.14 a. The limiting reactant when cooking with a gas grill would be the propane. This makes sense since propane is the material you must purchase in order to cook your food.

b. Since the chemical reaction only requires propane and oxygen, if the grill will not light with ample propane present, the limiting reactant must be the oxygen.

c. Once again, here is a case where you have adequate propane, so you can conclude that a yellow flame indicates that not enough oxygen is present to combust all of the propane. If there is not enough O_2 available for complete combustion, a reasonable assumption is that some of the products will have fewer oxygen atoms than CO_2. Therefore, a mixture of products would be obtained in this case, including carbon monoxide (CO) and soot (carbon particles).

3.15 a. This answer is unreasonable because 1.0×10^{-3} g is too small a weight for 0.33 mole of an element. For example, 0.33 mol of hydrogen, the lightest element, would have a mass of 0.33 g.

b. This answer is unreasonable because 1.80×10^{-10} g is too large for one water molecule. (The mass of one water molecule is 2.99×10^{-23} g)

c. This answer is reasonable because 3.01×10^{23} is one-half of Avogadro's number.

d. This answer is unreasonable because the units for molar mass should be g/mol, so this quantity is 1000 times too large.

3.16 a. In order to have a complete reaction, a ratio of two moles of hydrogen to every mole of oxygen is required. In this case, there is not enough oxygen in the air outside of the bubble for the complete reaction of hydrogen.

b. In this case, you have a ratio of one mole of H_2 to one mole of O_2. According to the balanced chemical reaction, every one mole of O_2 can react with two moles of H_2. In this case, when 0.5 mole of O_2 has reacted, all of the H_2 (one mole) will be consumed, leaving behind 0.5 mole of unreacted O_2.

(continued)

c. In this case, you have a ratio of two moles of H_2 to one mole of O_2, which is the correct stoichiometric amount, so all of the hydrogen and all of the oxygen react completely.

d. In order for reaction to occur, both oxygen and hydrogen must be present. Oxygen does not combust, and there is no hydrogen present to burn, so no reaction occurs.

3.17 a. The limiting reactant would be the charcoal because the air would supply as much oxygen as needed.

b. The limiting reactant would be the magnesium because the beaker would contain much more water than is needed for the reaction (approximately 18 mL of water is one mole).

c. The limiting reactant would be the H_2 because the air could supply as much nitrogen as is needed.

3.18 a. Since the balanced chemical equation for the reaction is $2H_2 + O_2 \rightarrow 2H_2O$ in order to form the water, you need two molecules of hydrogen for every one molecule of oxygen. Given the quantities of reactants present in the container and applying the 2:1 ratio, you can produce a maximum of twelve molecules of water.

b. The drawing of the container after the reaction should contain twelve H_2O molecules and two O_2 molecules.

3.19 a. The problem is that Avogadro's number was inadvertently used for the molar mass of calcium, which should be 40.08 g/mol. The correct calculation is

$$27.0 \text{ g Ca} \times \frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} = 0.67\underline{3}6 = 0.674 \text{ mol Ca}$$

b. The problem here is an incorrect mole ratio. There are two mol K^+ ions per one mol K_2SO_4. The correct calculation is

$$2.5 \text{ mol } K_2SO_4 \times \frac{2 \text{ mol } K^+ \text{ ions}}{1 \text{ mol } K_2SO_4} \times \frac{6.022 \times 10^{23} \text{ } K^+ \text{ ions}}{1 \text{ mol } K^+ \text{ ions}}$$

$$= 3.\underline{0}1 \times 10^{24} = 3.0 \times 10^{24} \text{ } K^+ \text{ ions}$$

c. The problem here is an incorrect mole ratio. The result should be

$$0.50 \text{ mol Na} \times \frac{2 \text{ mol } H_2O}{2 \text{ mol Na}} = 0.50 \text{ mol } H_2O$$

3.20 a. The missing concept is the mole ratio. His reasoning would only be correct if the reactants reacted in a one-to-one mole ratio. Here, the ratio is 5 mol O_2/2 mol C_2H_2. This means that, for every two moles of C_2H_2, five moles of O_2 are required. Since there are only four moles here, there is insufficient O_2, and it is the limiting reactant.

b. The missing concept again is the mole ratio. Since in this problem the reactants react in a one-to-one mole ratio, the reactant with the fewest moles is the limiting reactant, and the lucky guess works.

■ Solutions to Practice Problems

Note on significant figures: If the final answer to a solution needs to be rounded off, it is given first with one nonsignificant figure, and the last significant figure is underlined. The final answer is then rounded to the correct number of significant figures. In multiple-step problems, intermediate answers are given with at least one nonsignificant figure; however, only the final answer has been rounded off.

3.21 a. Formula weight of CH_3OH = AW of C + 4(AW of H) + AW of O. Using the values of atomic weights in the periodic table (inside front cover) rounded to four significant figures and rounding the answer to three significant figures, we have

$$\text{FW} = 12.01 \text{ amu} + (4 \times 1.008 \text{ amu}) + 16.00 \text{ amu} = 32.0\underline{4}2 = 32.0 \text{ amu}$$

b. FW of NO_3 = AW of N + 3(AW of O)

$$= 14.01 \text{ amu} + (3 \times 16.00 \text{ amu}) = 62.01 = 62.0 \text{ amu}$$

c. FW of K_2CO_3 = 2(AW of K) + AW of C + 3(AW of O)

$$= (2 \times 39.10 \text{ amu}) + 12.01 \text{ amu} + (3 \times 16.00 \text{ amu})$$

$$= 138.2\underline{1}0 = 138 \text{ amu}$$

d. FW of $Ni_3(PO_4)_2$ = 3(AW of Ni) + 2(AW of P) + 8(AW of O)

$$= (3 \times 58.70 \text{ amu}) + (2 \times 30.97 \text{ amu}) + (8 \times 16.00 \text{ amu})$$

$$= 366.0\underline{4}0 = 366 \text{ amu}$$

3.22 a. FW of $HC_2H_3O_2$ = 2(AW of C) + 4(AW of H) + 2(AW of O)

$$= (2 \times 12.01) + (4 \times 1.008) + (2 \times 16.00)$$

$$= 60.0\underline{5}2 = 60.1 \text{ amu}$$

(continued)

b. FW of PCl_5 = AW of P + 5(AW of Cl)
= 30.97 amu + (5 x 35.45 amu)
= $208.2\underline{2}0$ = 208 amu

c. FW of K_2SO_4 = 2(AW of K) + AW of S + 4(AW of O)
= (2 x 39.10 amu) + 32.07 amu + (4 x 16.00 amu)
= $174.2\underline{7}0$ = 174 amu (3 s.f.)

d. FW of $Ca(OH)_2$ = AW of Ca + 2(AW of H) + 2(AW of O)
= 40.08 amu + (2 x 1.008 amu) + (2 x 16.00 amu)
= $74.0\underline{9}6$ = 74.1 amu (3 s.f.)

3.23 a	SO_2	1 x AW of S			=	32.07 amu
		2 x AW of O	=	2 x 16.00	=	$\underline{32.00\text{ amu}}$
				MW of NO_2	=	$64.\underline{0}7$ amu = 64.1 amu (3 s.f.)
b.	PCl_3	1 x AW of P			=	30.97 amu
		3 x AW of Cl	=	3 x 35.45	=	$\underline{106.35\text{ amu}}$
				MW of PCl_3	=	137.32 amu = 137 amu (3 s.f.)

3.24 a	HNO_2	1 x AW of H			=	1.008 amu
		1 x AW of N			=	14.01 amu
		2 x AW of O	=	2 x 16.00	=	$\underline{32.00\text{ amu}}$
				MW of HNO_2	=	$47.\underline{0}18$ amu = 47.0 amu (3 s.f.)
b.	CO	1 x AW of C			=	12.01 amu
		1 x AW of O			=	$\underline{16.00\text{ amu}}$
				MW of CO	=	$28.\underline{0}1$ amu = 28.0 amu (3 s.f.)

3.25 First, find the formula weight of NH_4NO_3 by adding the respective atomic weights. Then convert it to the molar mass:

FW of NH_4NO_3 = 2(AW of N) + 4(AW of H) + 3(AW of O)
= (2 x 14.01 amu) + (4 x 1.008 amu) + (3 x 16.00 amu)
= $80.0\underline{5}2$ amu

The molar mass of NH_4NO_3 = 80.05 g/mol.

3.26 First, find the formula weight of H_3PO_4 by adding the respective atomic weights. Then convert it to the molar mass:

$$\begin{aligned}\text{FW of } H_3PO_4 &= 3(\text{AW of H}) + \text{AW of P} + 4(\text{AW of O})\\ &= (3 \times 1.008) + 30.97 + (4 \times 16.00)\\ &= 97.9\underline{9}4 \text{ amu}\end{aligned}$$

The molar mass of H_3PO_4 = 97.99 g/mol.

3.27 a. The atomic weight of Na equals 22.99 amu; thus, the molar mass equals 22.99 g/mol. Because one mol of Na atoms equals 6.022×10^{23} Na atoms, we calculate

$$\text{Mass of one Na atom} = \frac{22.99 \text{ g/mol}}{6.022 \times 10^{23} \text{ atom/mol}} = 3.81\underline{7}7 \times 10^{-23} \text{ g/atom}$$

b. The atomic weight of N equals 14.01 amu; thus, the molar mass equals 14.01 g/mol. Because one mol of N atoms equals 6.022×10^{23} N atoms, we calculate

$$\text{Mass of one N atom} = \frac{14.01 \text{ g/mol}}{6.022 \times 10^{23} \text{ atom/mol}} = 2.32\underline{6}4 \times 10^{-23} \text{ g/atom}$$

c. The formula weight of CH_3Cl = [12.01 + (3 x 1.008) + 35.45] = 50.48 amu; thus, the molar mass equals 50.48 g/mol. Because one mol of CH_3Cl molecules equals 6.022×10^{23} CH_3Cl molecules, we calculate

$$\text{Mass of one } CH_3Cl \text{ molecule} = \frac{50.48 \text{ g/mol}}{6.022 \times 10^{23} \text{ molecules/mol}}$$

$$= 8.38\underline{2}6 \times 10^{-23} \text{ g/molecule}$$

d. The formula weight of $Hg(NO_3)_2$ = 200.59 + (2 x 14.01) + (6 x 16.00)] = 324.61 amu; thus, the molar mass equals 324.61 g/mol. Because one formula weight of $Hg(NO_3)_2$ equals 6.022×10^{23} $Hg(NO_3)_2$ formula units, we calculate

$$\text{Mass of one } Hg(NO_3)_2 = \frac{324.61 \text{ g/mol}}{6.022 \times 10^{23} \text{ units/mol}} = 5.39\underline{0}4 \times 10^{-22} \text{ g/unit}$$

3.28 As in the previous problem, the atomic weights are used with units of g/mol. In addition, the formula weights have been found by addition of the atomic weights and are expressed in g/mol.

a. $$\text{Mass of one Fe atom} = \frac{55.85 \text{ g/mol}}{6.022 \times 10^{23} \text{ atom/mol}}\ 9.27\underline{4}3 \times 10^{-23} \text{ g/atom}$$

(continued)

b. Mass of one F atom $= \dfrac{19.00 \text{ g/mol}}{6.022 \times 10^{23} \text{ atom/mol}} = 3.15\underline{5}1 \times 10^{-23}$ g/atom

c. Mass of one N_2O molecule $= \dfrac{44.02 \text{ g/mol}}{6.022 \times 10^{23} \text{ molecules/mol}}$

$= 7.30\underline{9}9 \times 10^{-23}$ g/molecule

d. Mass of one $Al(OH)_3$ unit $= \dfrac{78.00 \text{ g/mol}}{6.022 \times 10^{23} \text{ unit/mol}} = 1.29\underline{5}3 \times 10^{-22}$ g/unit

3.29 First, find the formula weight (in amu) using the periodic table (inside front cover):

FW of $(CH_3CH_2)_2O$ = (4 x 12.01 amu) + (10 x 1.008 amu) + 16.00 amu
= 74.12 amu

Mass of one $(CH_3CH_2)_2O$ molecule $= \dfrac{74.12 \text{ g/mol}}{6.022 \times 10^{23} \text{ molecules/mol}}$

$= 1.23\underline{0}8 \times 10^{-22}$ g/molecule

3.30 First, find the formula weight (in amu) using the periodic table (inside front cover):

FW of glycerol = (3 x 12.01 amu) + (8 x 1.008 amu) + (3 x 16.00 amu)

= 92.09 amu

Mass of one glycerol molecule $= \dfrac{92.09 \text{ g/mol}}{6.022 \times 10^{23} \text{ molecules/mol}}$

$= 1.5\underline{2}9 \times 10^{-22}$ g/molecule

3.31 From the table of atomic weights, we obtain the following molar masses for parts a through d: Na = 22.99 g/mol; S = 32.07 g/mol; C = 12.01 g/mol; H = 1.008 g/mol; Cl = 35.45 g/mol; and N = 14.01 g/mol.

a. 0.15 mol Na x $\dfrac{22.99 \text{ g}}{1 \text{ mol Na}} = 3.\underline{4}48 = 3.4$ g Na

b. 0.594 mol S x $\dfrac{32.07 \text{ g S}}{1 \text{ mol S}} = 19.\underline{0}4 = 19.0$ g S

(continued)

c. Using molar mass = 84.93 g/mol for CH_2Cl_2, we obtain

$$2.78 \text{ mol } CH_2Cl_2 \times \frac{84.93 \text{ g } CH_2Cl_2}{1 \text{ mol } CH_2Cl_2} = 23\underline{6}.1 = 236 \text{ g } CH_2Cl_2$$

d. Using molar mass = 68.14 g/mol for $(NH_4)_2S$, we obtain

$$38 \text{ mol } (NH_4)_2S \times \frac{68.14 \text{ g } (NH_4)_2S}{1 \text{ mol } (NH_4)_2S} = 2.\underline{5}8 \times 10^3 = 2.6 \times 10^3 \text{ g } (NH_4)_2S$$

3.32 From the table of atomic weights, we obtain the following molar masses for parts a through d: C = 12.01 g/mol; O = 16.00 g/mol; K = 39.10 g/mol; Cr = 52.00 g/mol; Fe = 55.85 g/mol; and F = 19.00 g/mol.

a. $0.205 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 11.\underline{4}4 = 11.4 \text{ g Fe}$

b. $0.79 \text{ mol F} \times \frac{19.00 \text{ g F}}{1 \text{ mol F}} = 1\underline{5}.01 = 15 \text{ g F}$

c. Using molar mass = 44.01 g/mol for CO_2, we obtain

$$5.8 \text{ mol } CO_2 \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 2\underline{5}5.2 = 2.6 \times 10^2 \text{ g } CO_2$$

d. Using molar mass 194.20 g/mol for K_2CrO_4, we obtain

$$48.1 \text{ mol } K_2CrO_4 \times \frac{194.20 \text{ g } K_2CrO_4}{1 \text{ mol } K_2CrO_4} = 93\underline{4}1.02 = 9.34 \times 10^3 \text{ g } K_2CrO_4$$

3.33 First, find the molar mass of H_3BO_3: (3 x 1.008 amu) + 10.81 amu + (3 x 16.00 amu) = 61.83. Therefore, the molar mass of H_3BO_3 = 61.83 g/mol. The mass of H_3BO_3 is calculated as follows:

$$0.543 \text{ mol } H_3BO_3 \times \frac{61.83 \text{ g } H_3BO_3}{1 \text{ mol } H_3BO_3} = 33.\underline{5}7 = 33.6 \text{ g } H_3BO_3$$

3.34 First, find the molar mass of CS_2: 12.01 amu + (2 x 32.07 amu) = 76.15 amu. Therefore, the molar mass of CS_2 = 76.15 g/mol. The mass of CS_2 is calculated as follows:

$$0.0205 \text{ mol } CS_2 \times \frac{76.15 \text{ g } CS_2}{1 \text{ mol } CS_2} = 1.5\underline{6}10 = 1.56 \text{ g } CS_2$$

3.35 From the table of atomic weights, we obtain the following rounded molar masses for parts a through d: C = 12.01 g/mol; Cl = 35.45 g/mol; H = 1.008 g/mol; Al = 26.98 g/mol; and O = 16.00 g/mol.

a. $2.86 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.23\underline{8}1 = 0.238 \text{ mol C}$

b. $7.05 \text{ g } Cl_2 \times \frac{1 \text{ mol } Cl_2}{70.90 \text{ g } Cl_2} = 0.099\underline{4}3 = 0.0994 \text{ mol } Cl_2$

c. The molar mass of C_4H_{10} = (4 x 12.01) + (10 x 1.008) = 58.12 g C_4H_{10}/mol C_4H_{10}. The mass of C_4H_{10} is calculated as follows:

$$76 \text{ g } C_4H_{10} \times \frac{1 \text{ mol } C_4H_{10}}{58.12 \text{ g } C_4H_{10}} = 1.\underline{3}07 = 1.3 \text{ mol } C_4H_{10}$$

d. The molar mass of $Al_2(CO_3)_3$ = (2 x 26.98) + (3 x 12.01) + (9 x 16.00 g) = 233.99 g/mol $Al_2(CO_3)_3$. The mass of $Al_2(CO_3)_3$ is calculated as follows:

$$26.2 \text{ g } Al_2(CO_3)_3 \times \frac{1 \text{ mol } Al_2(CO_3)_3}{233.99 \text{ g } Al_2(CO_3)_3} = 0.11\underline{1}9 = 0.112 \text{ mol } Al_2(CO_3)_3$$

3.36 From the table of atomic weights, we obtain the following rounded molar masses for parts a through d: As = 74.92 g/mol; S = 32.07 g/mol; N = 14.01 g/mol; H = 1.008 g/mol; Al = 26.98 g/mol; and O = 16.00 g/mol.

a. $2.57 \text{ g As} \times \frac{1 \text{ mol As}}{74.92 \text{ g As}} = 0.034\underline{3}0 = 0.0343 \text{ mol As}$

b. $7.83 \text{ g } S_8 \times \frac{1 \text{ mol } S_8}{256.56 \text{ g } S_8} = 0.030\underline{5}1 = 0.0305 \text{ mol } S_8$

c. The molar mass of N_2H_4 = (2 x 14.01) + (4 x 1.008) = 32.052 g N_2H_4/mol N_2H_4. The mass of N_2H_4 is calculated as follows:

$$36.5 \text{ g} \times \frac{1 \text{ mol } N_2H_4}{32.052 \text{ g } N_2H_4} = 1.1\underline{3}87 = 1.14 \text{ g}$$

d. The molar mass of $Al_2(SO_4)_3$ = (2 x 26.98) + (3 x 32.07) + (12 x 16.00) = 342.17 g $Al_2(SO_4)_3$/mol $Al_2(SO_4)_3$. The mass of $Al_2(SO_4)_3$ is calculated as follows:

$$227 \text{ g } Al_2(SO_4)_3 \times \frac{1 \text{ mol } Al_2(SO_4)_3}{342.17 \text{ g } Al_2(SO_4)_3} \; 0.66\underline{3}4 = 0.663 \text{ mol } Al_2(SO_4)_3$$

3.37 Calculate the formula weight of calcium sulfate: 40.08 amu + 32.07 amu + (4 x 16.00 amu) = 136.15 amu. Therefore, the molar mass of $CaSO_4$ is 136.15 g/mol. Use this to convert the mass of $CaSO_4$ to moles:

$$0.791\,\text{g CaSO}_4 \times \frac{1\,\text{mol CaSO}_4}{136.15\,\text{g CaSO}_4} = 5.8\underline{1}1 \times 10^{-3} = 5.81 \times 10^{-3}\ \text{mol CaSO}_4$$

Calculate the molecular weight of water: (2 x 1.008 amu) + 16.00 amu = 18.02 amu. Therefore, the molar mass of H_2O equals 18.02 g/mol. Use this to convert the rest of the sample to moles of water:

$$0.209\,\text{g H}_2\text{O} \times \frac{1\,\text{mol H}_2\text{O}}{18.02\,\text{g H}_2\text{O}} = 1.1\underline{5}9 \times 10^{-2} = 1.16 \times 10^{-2}\ \text{mol H}_2\text{O}$$

Because 0.01159 mol is about twice 0.005811 mol, both numbers of moles are consistent with the formula, $CaSO_4{\bullet}2H_2O$.

3.38 Calculate the formula weight of copper(II) sulfate: 63.55 amu + 32.07 amu + (4 x 16.00 amu) = 159.62 amu. Thus, the molar mass of $CuSO_4$ is 159.62 g/mol. From the previous problem, the molar mass of H_2O is 18.02 g/mol. Use this to convert the rest of the sample to moles of water:

$$0.558\,\text{g H}_2\text{O} \times \frac{1\,\text{mol H}_2\text{O}}{18.02\,\text{g H}_2\text{O}} = 0.030\underline{9}6\ \text{mol H}_2\text{O}$$

Calculate the moles of $CuSO_4$ to be able to compare relative molar amounts of $CuSO_4$ and H_2O. Then, divide the moles of H_2O by the moles of $CuSO_4$:

$$0.989\,\text{g CuSO}_4 \times \frac{1\,\text{mol CuSO}_4}{159.62\,\text{g CuSO}_4} = 0.0061\underline{9}6\ \text{mol CuSO}_4$$

0.03096 mol H_2O ÷ 0.006196 mol $CuSO_4$ = 4.99/1, or about 5:1

(consistent with $CuSO_4{\bullet}5H_2O$)

3.39 The following rounded atomic weights are used: Li = 6.94 g/mol; Br = 79.90 g/mol; N = 14.01 g/mol; H = 1.008 g/mol; Pb = 207.2 g/mol; Cr = 52.00 g/mol; O = 16.00 g/mol; and S = 32.07 g/mol. Also, Avogadro's number is 6.022×10^{23} atoms, so

a. $$\text{No. Li atoms} = 8.21\,\text{g Li} \times \frac{6.022 \times 10^{23}\ \text{atoms}}{6.941\,\text{g Li}} = 7.1\underline{2}2 \times 10^{23}\ \text{atoms}$$

b. $$\text{No. Br atoms} = 32.0\,\text{g Br}_2 \times \frac{2 \times 6.022 \times 10^{23}\ \text{atoms}}{(2 \times 79.90)\,\text{g Br}_2} = 2.4\underline{1}2 \times 10^{23}\ \text{atoms}$$

(continued)

c. $\text{No. NH}_3\text{ molecules} = 45\text{ g NH}_3 \times \dfrac{6.022 \times 10^{23}\text{ molecules}}{17.03\text{ g NH}_3} = 1.\underline{5}9 \times 10^{24}\text{ molecules}$

d. $\text{No. PbCrO}_4\text{ units} = 201\text{ g PbCrO}_4 \times \dfrac{6.022 \times 10^{23}\text{ units}}{323.2\text{ g PbCrO}_4} = 3.7\underline{4}5 \times 10^{23}\text{ units}$

e. $\text{No. SO}_4^{2-}\text{ ions} = 14.3\text{ g Cr}_2(\text{SO}_4)_3 \times \dfrac{3 \times 6.022 \times 10^{23}\text{ ions}}{392.21\text{ g Cr}_2(\text{SO}_4)_3} = 6.5\underline{8}7 \times 10^{22}\text{ ions}$

3.40 These rounded atomic weights are used: Al = 26.98 g/mol; I = 126.90 g/mol; N = 14.01 g/mol; O = 16.00 g/mol; Na = 22.99 g/mol; Cl = 35.45 g/mol; Ca = 40.08 g/mol; and P = 30.97 g/mol. Also, Avogadro's number is 6.022×10^{23} atoms, so

a. $\text{No. Al atoms} = 25.7\text{ g Al} \times \dfrac{6.022 \times 10^{23}\text{ atoms}}{26.98\text{ g Al}} = 5.7\underline{3}6 \times 10^{23}\text{ atoms}$

b. $\text{No. I atoms} = 8.71\text{ g I}_2 \times \dfrac{2 \times 6.022 \times 10^{23}\text{ atoms}}{2 \times 126.90\text{ g I}_2} = 4.1\underline{3}3 \times 10^{22}\text{ atoms}$

c. $\text{No. N}_2\text{O}_5\text{ molecules} = 14.9\text{ g N}_2\text{O}_5 \times \dfrac{6.022 \times 10^{23}\text{ molecules}}{108.02\text{ g N}_2\text{O}_5}$

$= 8.3\underline{0}6 \times 10^{22}\text{ molecules}$

d. $\text{No. NaClO}_4\text{ units} = 3.31\text{ g NaClO}_4 \times \dfrac{6.022 \times 10^{23}\text{ units}}{122.44\text{ g NaClO}_4} = 1.6\underline{2}8 \times 10^{22}\text{ units}$

e. $\text{No. Ca}^{2+}\text{ ions} = 4.71\text{ g Ca}_3(\text{PO}_4)_2 \times \dfrac{3 \times 6.022 \times 10^{23}\text{ ions}}{310.18\text{ g Ca}_3(\text{PO}_4)_2} = 2.7\underline{4}3 \times 10^{22}\text{ ions}$

3.41 Calculate the molecular weight of CCl_4: 12.01 amu + (4 x 35.45 amu) = 153.81 amu. Use this and Avogadro's number to express it as 153.81 g/N_A to calculate the number of molecules:

$$7.58\text{ mg CCl}_4 \times \frac{1\text{ g}}{1000\text{ mg}} \times \frac{6.022 \times 10^{23}\text{ molecules}}{153.81\text{ g CCl}_4} = 2.9\underline{6}8 \times 10^{19}$$

$$= 2.97 \times 10^{19}\text{ molecules}$$

3.42 Calculate the molecular weight of ClF_3: 35.45 amu + (3 x 19.00 amu) = 92.45 amu. Use this and Avogadro's number to express it as 92.45 g/N_A to calculate the number of molecules:

$$5.88 \text{ mg } ClF_3 \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{92.45 \text{ g } ClF_3} = 3.8\underline{3}0 \times 10^{19}$$

$$= 3.83 \times 10^{19} \text{ molecules}$$

3.43 Mass percentage carbon $= \frac{\text{mass of C in sample}}{\text{mass of sample}} \times 100\%$

$$\text{Percent carbon} = \frac{1.584 \text{ g}}{1.836 \text{ g}} \times 100\% = 86.2\underline{7}4 = 86.27\%$$

3.44 Mass percentage alcohol $= \frac{\text{mass of alcohol in solution}}{\text{mass of sample}} \times 100\%$

$$\text{Percent alcohol} = \frac{3.98 \text{ g}}{6.01 \text{ g}} \times 100\% = 66.\underline{2}2 = 66.2\%$$

3.45 Mass percentage phosphorus oxychloride $= \frac{\text{mass of } POCl_3 \text{ in sample}}{\text{mass of sample}} \times 100\%$

$$\text{Percent } POCl_3 = \frac{1.72 \text{ mg}}{8.53 \text{ mg}} \times 100\% = 20.\underline{1}6 = 20.2\%$$

3.46 Mass percentage sulfur $= \frac{\text{mass of S in sample}}{\text{mass of sample}} \times 100\%$

$$\text{Percent sulfur} = \frac{1.64 \text{ mg}}{3.17 \text{ mg}} \times 100\% = 51.\underline{7}3 = 51.7\%$$

3.47 Start with the definition for percentage nitrogen, and rearrange this equation to find the mass of N in the fertilizer.

$$\text{Mass percentage nitrogen} = \frac{\text{mass of N in fertilizer}}{\text{mass of fertilizer}} \times 100\%$$

$$\text{Mass N} = \frac{\text{mass \% N}}{100\%} \times \text{mass of fertilizer} = \frac{14.0\%}{100\%} \times 4.15 \text{ kg} = 0.58\underline{1}0$$

$$= 0.581 \text{ kg N}$$

3.48 Start by finding the mass of 1.000 L of seawater using the density of 1.025 g/cm^3.

$$\text{Mass seawater} = 1.00 \text{ L} \times \frac{10^3 \text{ cm}^3}{1 \text{L}} \times \frac{1.025 \text{ g}}{1 \text{cm}^3} = 1.025 \times 10^3 \text{ g}$$

Continue with the definition for percentage of bromine in the seawater, and rearrange this equation to find the mass of Br in the seawater.

$$\text{Mass percentage Br} = \frac{\text{mass of Br in seawater}}{\text{mass of seawater}} \times 100\%$$

$$\text{Mass Br} = \frac{\text{mass \% Br}}{100\%} \times \text{mass seawater} = \frac{0.0065\%}{100\%} \times 1025 \text{ g} = 0.06\underline{6}6$$

$$= 0.067 \text{ g Br}$$

3.49 Convert moles to mass using the molar masses from the respective atomic weights. Then, calculate the mass percentages from the respective masses.

$$0.0898 \text{ mol Al} \times \frac{26.98 \text{ g Al}}{1 \text{mol Al}} = 2.4\underline{2}2 \text{ g Al}$$

$$0.0381 \text{ mol Mg} \times \frac{24.31 \text{g Mg}}{1 \text{mol Mg}} = 0.92\underline{6}2 \text{ g Mg}$$

$$\text{Percent Al} = \frac{\text{mass of Al}}{\text{mass of alloy}} = \frac{2.422 \text{ g Al}}{3.349 \text{ g alloy}} \times 100\% = 72.\underline{3}4 = 72.3\% \text{ Al}$$

$$\text{Percent Mg} = \frac{\text{mass of Mg}}{\text{mass of alloy}} = \frac{0.9262 \text{ g Mg}}{3.349 \text{ g alloy}} \times 100\% = 27.\underline{6}55 = 27.7\% \text{ Mg}$$

3.50 Convert moles to mass using the molar masses from the respective atomic weights of 20.18 g/mol for Ne and 83.80 g/mol for Kr. Then, calculate the mass percentages from the respective masses.

$$0.0856 \text{ mol Ne} \times \frac{20.18 \text{ g Ne}}{1 \text{ mol Ne}} = 1.7\underline{2}7 \text{ g Ne}$$

$$0.0254 \text{ mol Kr} \times \frac{83.80 \text{ g Kr}}{1 \text{ mol Kr}} = 2.1\underline{2}9 \text{ g Kr}$$

$$\text{Percent Ne} = \frac{\text{mass of Ne}}{\text{mass of mix}} = \frac{1.727 \text{ g Ne}}{3.856 \text{ g mix}} \times 100\% = 44.\underline{7}9 = 44.8\% \text{ Ne}$$

$$\text{Percent Kr} = \frac{\text{mass of Kr}}{\text{mass of mix}} = \frac{2.129 \text{ g Kr}}{3.856 \text{ g mix}} \times 100\% = 55.\underline{2}1 = 55.2\% \text{ Kr}$$

3.51 In each part, the numerator consists of the mass of the element in one mole of the compound; the denominator is the mass of one mole of the compound. Use the atomic weights of C = 12.01 g/mol; O = 16.00 g/mol; Na = 22.99 g/mol; H = 1.008 g/mol; P = 30.97 g/mol; Co = 58.93 g/mol; and N = 14.01 g/mol.

a. $$\text{Percent C} = \frac{\text{mass of C}}{\text{mass of CO}} = \frac{12.01 \text{ g C}}{28.01 \text{ g CO}} \times 100\% = 42.878 = 42.9\%$$

$$\text{Percent O} = 100.000\% - 42.8\underline{7}8\% \text{C} = 57.122 = 57.1\%$$

b. $$\text{Percent C} = \frac{\text{mass of C}}{\text{mass of } CO_2} = \frac{12.01 \text{ g C}}{44.01 \text{ g } CO_2} \times 100\% = 27.\underline{2}89 = 27.3\%$$

$$\text{Percent O} = 100.000\% - 27.2\underline{8}9\% \text{ C} = 72.711 = 72.7\%$$

c. $$\text{Percent Na} = \frac{\text{mass of Na}}{\text{mass of } NaH_2PO_4} = \frac{22.99 \text{ g Na}}{119.98 \text{ g } NaH_2PO_4} \times 100\%$$

$$= 19.161 = 19.2\%$$

$$\text{Percent H} = \frac{\text{mass of H}}{\text{mass of } NaH_2PO_4} = \frac{2.016 \text{ g H}}{119.98 \text{ g } NaH_2PO_4} \times 100\% = 1.6802 = 1.68\%$$

$$\text{Percent P} = \frac{\text{mass of P}}{\text{mass of } NaH_2PO_4} = \frac{30.97 \text{ g P}}{119.98 \text{ g } NaH_2PO_4} \times 100\% = 25.812 = 25.8\%$$

$$\text{Percent O} = 100.000\% - (19.161 + 1.6802 + 25.812) = 53.346 = 53.3\%$$

(continued)

d. $\text{Percent Co} = \dfrac{\text{mass of Co}}{\text{mass of Co(NO}_3)_2} = \dfrac{58.93 \text{ g Co}}{182.95 \text{ g Co(NO}_3)_2} \times 100\%$

$= 32.211 = 32.2\%$

$\text{Percent N} = \dfrac{\text{mass of N}}{\text{mass of Co(NO}_3)_2} = \dfrac{2 \times 14.01 \text{ g N}}{182.95 \text{ g Co(NO}_3)_2} \times 100\%$

$= 15.316 = 15.3\%$

$\text{Percent O} = 100.000\% - (32.2\underline{1}1 + 15.3\underline{1}6) = 52.473 = 52.5\%$

3.52 In each part, the numerator consists of the mass of the element in one mole of the compound; the denominator is the mass of one mole of the compound

a. $\text{Percent N} = \dfrac{\text{mass of N}}{\text{mass of NO}} = \dfrac{14.01 \text{ g N}}{30.01 \text{ g NO}} \times 100\% = 46.684 = 46.7\%$

$\text{Percent O} = 100.000\% - 46.684\% \text{ N} = 53.316 = 53.3\%$

b. $\text{Percent H} = \dfrac{\text{mass of H}}{\text{mass of H}_2\text{O}_2} = \dfrac{2.016 \text{ g H}}{34.02 \text{ g H}_2\text{O}_2} \times 100\% = 5.92\underline{5}9 = 5.93\%$

$\text{Percent O} = 100.000\% - 5.2959\% \text{ H} = 94.07 = 94.1\%$

c. $\text{Percent K} = \dfrac{\text{mass of K}}{\text{mass of KClO}_4} = \dfrac{39.10 \text{ g K}}{138.55 \text{ g KClO}_4} \times 100\% = 28.221 = 28.2\%$

$\text{Percent Cl} = \dfrac{\text{mass of Cl}}{\text{mass of KClO}_4} = \dfrac{35.45 \text{ g Cl}}{138.55 \text{ g KClO}_4} \times 100\% = 25.586 = 25.6\%$

$\text{Percent O} = 100.000\% - (28.221 + 25.586) = 46.193 = 46.2\%$

d. $\text{Percent Mn} = \dfrac{\text{mass of Mn}}{\text{mass of Mn(NO}_2)_2} = \dfrac{54.94 \text{ g Mn}}{146.96 \text{ g Mn(NO}_2)_2} = 37.384 = 37.4\%$

$\text{Percent N} = \dfrac{\text{mass of N}}{\text{mass of Mn(NO}_2)_2} \; \dfrac{2 \times 14.01 \text{ g N}}{146.96 \text{ g Mn(NO}_2)_2} = 19.066 = 19.1\%$

$\text{Percent O} = 100.000\% - (37.384 + 19.066) = 43.550 = 43.5\%$

3.53 The molecular model of toluene contains seven carbon atoms and eight hydrogen atoms, so the molecular formula of toluene is C_7H_8. The molar mass of toluene is 92.134 g/mol. The mass percentages are

$$\text{Percent C} = \frac{\text{mass of C}}{\text{mass of } C_7H_8} = \frac{7 \times 12.01 \text{ g}}{92.134 \text{ g}} \times 100\% = 91.247 = 91.2\%$$

$$\text{Percent H} = 100\% - 91.247 = 8.753 = 8.75\%$$

3.54 The molecular model of 2-propanol contains three carbon atoms, eight hydrogen atoms, and one oxygen atom, so the molecular formula of 2-propanol is C_3H_8O. The molar mass of 2-propanol is 60.094 g/mol. The mass percentages are

$$\text{Percent C} = \frac{\text{mass of C}}{\text{mass of } C_3H_8O} = \frac{3 \times 12.01 \text{ g}}{60.094 \text{ g}} \times 100\% = 59.956 = 60.0\%$$

$$\text{Percent H} = \frac{\text{mass of H}}{\text{mass of } C_3H_8O} = \frac{8 \times 1.008 \text{ g}}{60.094 \text{ g}} \times 100\% = 13.418 = 13.4\%$$

$$\text{Percent O} = 100\% - (59.956 + 13.418) = 26.626 = 26.6\%$$

3.55 Find the moles of C in each amount in one-step operations. Calculate the moles of each compound using the molar mass; then multiply by the number of moles of C per mole of compound:

$$\text{Mol C (glucose)} = 6.01 \text{ g} \times \frac{1 \text{ mol}}{180.2 \text{ g}} \times \frac{6 \text{ mol C}}{1 \text{ mol glucose}} = 0.20\underline{0} \text{ mol}$$

$$\text{Mol C (ethanol)} = 5.85 \text{ g} \times \frac{1 \text{ mol}}{46.07 \text{ g}} \times \frac{2 \text{ mol C}}{1 \text{ mol ethanol}} = 0.25\underline{4} \text{ mol (more C)}$$

3.56 Find the moles of S in each amount in one-step operations. Calculate the moles of each compound using the molar mass; then multiply by the number of moles of S per mole of compound.

$$\text{Mol S } (CaSO_4) = 40.8 \text{ g} \times \frac{1 \text{ mol } CaSO_4}{136.15 \text{ g}} \times \frac{1 \text{ mol S}}{1 \text{ mol } CaSO_4}$$

$$= 0.29\underline{9}7 \text{ mol (more S)}$$

$$\text{Mol S } (Na_2SO_3) = 35.2 \text{ g} \times \frac{1 \text{ mol } Na_2SO_3}{126.05 \text{ g}} \times \frac{1 \text{ mol S}}{1 \text{ mol } Na_2SO_3}$$

$$= 0.27\underline{9}3 \text{ mol}$$

3.57 First, calculate the mass of C in the glycol by multiplying the mass of CO_2 by the molar mass of C and the reciprocal of the molar mass of CO_2. Then, calculate the mass of H in the glycol by multiplying the mass of H_2O by the molar mass of 2H and the reciprocal of the molar mass of H_2O. Then, use the masses to calculate the mass percentages. Calculate O by difference.

$$9.06\text{ mg CO}_2 \times \frac{1\text{ mol CO}_2}{44.01\text{ g CO}_2} \times \frac{12.01\text{ g C}}{1\text{ mol C}} = 2.4\underline{7}2\text{ mg C}$$

$$5.58\text{ mg H}_2\text{O} \times \frac{1\text{ mol H}_2\text{O}}{18.02\text{ g H}_2\text{O}} \times \frac{2\text{ H}}{1\text{H}_2\text{O}} \times \frac{1.008\text{ g H}}{1\text{ mol H}} = 0.62\underline{4}3\text{ mg H}$$

Mass O = 6.38 mg - (2.472 + 0.6243) = 3.2$\underline{8}$4 mg O

Percent C = (2.472 mg C/6.38 mg glycol) x 100% = 38.$\underline{7}$4 = 38.7%

Percent H = (0.6243 mg H/6.38 mg glycol) x 100% = 9.7$\underline{8}$5 = 9.79%

Percent O = (3.284 mg O/6.38 mg glycol) x 100% = 51.$\underline{4}$7 = 51.5%

3.58 First, calculate the mass of C in the phenol by multiplying the mass of CO_2 by the molar mass of C and the reciprocal of the molar mass of CO_2. Then, calculate the mass of H in the phenol by multiplying the mass of H_2O by the molar mass of 2H and the reciprocal of the molar mass of H_2O. Then, use the masses to calculate the mass percentages. Calculate O by difference.

$$14.67\text{ mg CO}_2 \times \frac{1\text{ mol CO}_2}{44.01\text{ g CO}_2} \times \frac{12.01\text{ g C}}{1\text{ mol C}} = 4.00\underline{3}3\text{ mg C}$$

$$3.01\text{ mg H}_2\text{O} \times \frac{1\text{ mol H}_2\text{O}}{18.02\text{ g CO}_2} \times \frac{2\text{ H}}{1\text{H}_2\text{O}} \times \frac{1.008\text{ g H}}{1\text{ mol H}} = 0.33\underline{6}8\text{ mg H}$$

Mass O = 5.23 mg - (4.0033 + 0.3368) = 0.8$\underline{8}$99 mg O

Percent C = (4.0033 mg/5.23 mg) x 100% = 76.$\underline{5}$4 = 76.5%

Percent H = (0.3368 mg/5.23 mg) x 100% = 6.4$\underline{3}$9 = 6.44%

Percent O = (0.8899 mg/5.23 mg) x 100% = 1$\underline{7}$.0 = 17%

3.59 Start by calculating the moles of Os and O; then divide each by the smaller number of moles to obtain integers for the empirical formula.

$$\text{Mol Os} = 2.16 \text{ g Os} \times \frac{1 \text{ mol Os}}{190.2 \text{ g Os}} = 0.011\underline{3}6 \text{ mol (smaller number)}$$

$$\text{Mol O} = (2.89 - 2.16) \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.04\underline{5}6 \text{ mol}$$

$$\text{Integer for Os} = 0.01136 \div 0.01136 = 1.0\underline{0}0$$

$$\text{Integer for O} = 0.0456 \div 0.01136 = 4.\underline{0}1$$

Within experimental error, the empirical formula is OsO_4.

3.60 Start by calculating the moles of W and O; then divide each by the smaller number of moles to obtain integers for the empirical formula.

$$\text{Mol W} = 4.23 \text{ g W} \times \frac{1 \text{ mol W}}{183.85 \text{ g W}} = 0.023\underline{0}1 \text{ mol (smaller number)}$$

$$\text{Mol O} = (5.34 - 4.23) \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.069\underline{3}8 \text{ mol}$$

$$\text{Integer for W} = 0.02301 \div 0.02301 = 1.0\underline{0}0$$

$$\text{Integer for O} = 0.06938 \div 0.02301 = 3.0\underline{1}5$$

Because 3.015 = 3.0 within experimental error, the empirical formula is WO_3

3.61 Assume a sample of 100.0 g of potassium manganate. By multiplying this by the percentage composition, we obtain 39.6 g of K, 27.9 g of Mn, and 32.5 g of O. Convert each of these masses to moles by dividing by molar mass.

$$\text{Mol K} = 39.6 \text{ g K} \times \frac{1 \text{ mol K}}{39.10 \text{ g K}} = 1.0\underline{1}3 \text{ mol}$$

$$\text{Mol Mn} = 29.7 \text{ g Mn} \times \frac{1 \text{ mol Mn}}{54.94 \text{ g Mn}} = 0.50\underline{7}8 \text{ mol (smallest number)}$$

$$\text{Mol O} = 32.5 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.0\underline{3}1 \text{ mol}$$

(continued)

Now, divide each number of moles by the smallest number to obtain the smallest set of integers for the empirical formula.

Integer for K = 1.013 ÷ 0.5078 = 1.998, or 2

Integer for Mn = 0.5078 ÷ 0.5078 = 1.000, or 1

Integer for O = 2.031 ÷ 0.5078 = 3.999, or 4

The empirical formula is thus K_2MnO_4.

3.62 Assume a sample of 100.0 g of hydroquinone. By multiplying this by the percentage composition, we obtain 65.4 g of C, 5.5 g of H, and 29.1 g of O. Convert each of these masses to moles by dividing by the molar mass.

$$\text{Mol C} = 65.4 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 5.4\underline{4}5 \text{ mol}$$

$$\text{Mol H} = 5.5 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 5.\underline{4}6 \text{ mol}$$

$$\text{Mol O} = 29.1 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.8\underline{1}9 \text{ mol (smallest number)}$$

Now, divide each number of moles by the smallest number to obtain the smallest set of integers for the empirical formula.

Integer for C = 5.445 ÷ 1.819 = 2.99, or 3

Integer for H = 5.46 ÷ 1.819 = 3.0, or 3

Integer for O = 1.819 ÷ 1.819 = 1.00, or 1

The empirical formula is thus C_3H_3O.

3.63 Assume a sample of 100.0 g of acrylic acid. By multiplying this by the percentage composition, we obtain 50.0 g C, 5.6 g H, and 44.4 g O. Convert each of these masses to moles by dividing by the molar mass.

$$\text{Mol C} = 50.0 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.1\underline{6}3 \text{ mol}$$

$$\text{Mol H} = 5.6 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 5.\underline{5}6 \text{ mol}$$

(continued)

$$\text{Mol O} = 44.0 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.7\underline{7}5 \text{ mol (smallest number)}$$

Now, divide each number of moles by the smallest number to obtain the smallest number of moles and the tentative integers for the empirical formula.

Tentative integer for C = 4.163 ÷ 2.775 = 1.50, or 1.5

Tentative integer for H = 5.56 ÷ 2.775 = 2.00, or 2

Tentative integer for O = 2.775 ÷ 2.775 = 1.00, or 1

Because 1.5 is not a whole number, multiply each tentative integer by two to obtain the final integer for the empirical formula:

C: 2 x 1.5 = 3

H: 2 x 2 = 4

O: 2 x 1 = 2

The empirical formula is thus $C_3H_4O_2$.

3.64 Assume a sample of 100.0 g of malonic acid. By multiplying this by the percentage composition, we obtain 34.6 g C, 3.9 g H, and 61.5 g O. Convert each of these masses to moles by dividing by the molar mass.

$$\text{Mol C} = 34.6 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 2.8\underline{8}1 \text{ mol (smallest number)}$$

$$\text{Mol H} = 3.9 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 3.\underline{8}7 \text{ mol}$$

$$\text{Mol O} = 61.5 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.8\underline{4}4 \text{ mol}$$

Now, divide each number of moles by the smallest number to obtain the smallest number of moles and the tentative integers for the empirical formula.

Tentative integer for C = 2.881 ÷ 2.881 = 1.00, or 1

Tentative integer for H = 3.87 ÷ 2.881 = 1.34, or 1-1/3

Tentative integer for O = 3.884 ÷ 2.881 = 1.334, or 1-1/3

(continued)

Because 1-1/3 is not a whole number, multiply each tentative integer by three to give the final integer for the empirical formula:

C: 3 x 1 = 3

H: 3 x 1-1/3 = 4

O: 3 x 1-1/3 = 4

The empirical formula is thus $C_3H_4O_4$.

3.65 a. Assume for the calculation that you have 100.0 g; of this quantity, 92.25 g is C and 7.75 g is H. Now, convert these masses to moles:

$$92.25\text{ g C} \times \frac{1\text{ mol C}}{12.01\text{ g C}} = 7.68\underline{1}09\text{ mol C}$$

$$7.75\text{ g H} \times \frac{1\text{ mol H}}{1.008\text{ g H}} = 7.6\underline{8}8\text{ mol H}$$

Usually, you divide all the mole numbers by the smaller one, but in this case both are equal, so the ratio of the number of C atoms to the number of H atoms is 1:1. Thus, the empirical formula for both compounds is CH.

b. Obtain n, the number of empirical formula units in the molecule, by dividing the molecular weight of 52.03 amu and 78.05 amu by the empirical formula weight of 13.018 amu:

$$\text{For } 52.03: \ n = \frac{52.03\text{ amu}}{13.018\text{ amu}} = 3.99\underline{6}8\text{, or } 4$$

$$\text{For } 78.05: \ n = \frac{78.05\text{ amu}}{13.018\text{ amu}} = 5.99\underline{5}5\text{, or } 6$$

The molecular formulas are: for 52.03, $(CH)_4$ or C_4H_4; and for 78.05, $(CH)_6$ or C_6H_6.

3.66 a. Assume for the calculation that you have 100.0 g; of this quantity, 85.62 g is C and 14.38 g is H. Now convert these masses to moles:

$$85.62\text{ g C} = \times \frac{1\text{ mol C}}{12.01\text{ g C}} = 7.12\underline{9}05\text{ mol C}$$

$$14.38\text{ g H} \times \frac{1\text{ mol H}}{1.008\text{ g H}} = 14.\underline{2}6\text{ mol H}$$

(continued)

Divide both mole numbers by the smaller one:

$$\text{For C}: \frac{7.129\ \text{mol}}{7.129\ \text{mol}} = 1.0\underline{0}$$

$$\text{For H}: \frac{14.26\ \text{mol}}{7.129\ \text{mol}} = 2.0\underline{0}02$$

The empirical formula is obviously CH_2.

b. Obtain n, the number of empirical formula units in the molecule, by dividing the molecular weights of 28.03 amu and 56.06 amu by the empirical formula weight of 14.026 amu:

$$\text{For } 28.03: \ n = \frac{28.03\ \text{amu}}{14.026\ \text{amu}} = 1.99\underline{8}4, \text{ or } 2$$

$$\text{For } 56.06: \ n = \frac{56.06\ \text{amu}}{14.026\ \text{amu}} = 3.99\underline{6}8, \text{ or } 4$$

The molecular formulas are: for 28.03, $(CH_2)_2$ or C_2H_4; and for 56.06, $(CH_2)_4$ or C_4H_8.

3.67 The formula weight corresponding to the empirical formula C_2H_6N may be found by adding the respective atomic weights.

$$\begin{aligned}\text{Formula weight} &= (2 \times 12.01\ \text{amu}) + (6 \times 1.008\ \text{amu}) + 14.01\ \text{amu} \\ &= 44.08\ \text{amu}\end{aligned}$$

Dividing the molecular weight by the formula weight gives the number of times the C_2H_6N unit occurs in the molecule. Because the molecular weight is an average of 88.5 ([90 + 87] ÷ 2), this quotient is

$$88.5\ \text{amu} \div 44.1\ \text{amu} = 2.\underline{0}06, \text{ or } 2$$

Therefore, the molecular formula is $(C_2H_6N)_2$, or $C_4H_{12}N_2$.

3.68 The formula weight corresponding to the empirical formula BH_3 may be found by adding the respective atomic weights.

$$\text{Formula weight} = 10.81\ \text{amu} + (3 \times 1.008\ \text{amu}) = 13.83\ \text{amu}$$

Dividing the molecular weight by the formula weight gives the number of times the BH_3 unit occurs in the molecule. Because the molecular weight is 28 amu, this quotient is

$$28\ \text{amu} \div 13.83\ \text{amu} = 2.\underline{0}2$$

Therefore, the molecular formula is $(BH_3)_2$, or B_2H_6.

3.69 Assume a sample of 100.0 g of oxalic acid. By multiplying this by the percentage composition, we obtain 26.7 g C, 2.2 g H, and 71.1 g O. Convert each of these masses to moles by dividing by the molar mass.

$$\text{Mol C} = 26.7 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 2.2\underline{2}3 \text{ mol}$$

$$\text{Mol H} = 2.2 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 2.\underline{1}8 \text{ mol (smallest number)}$$

$$\text{Mol O} = 71.1 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 4.4\underline{4}3 \text{ mol}$$

Now, divide each number of moles by the smallest number to obtain the smallest set of integers for the empirical formula.

Integer for C = 2.223 ÷ 2.18 = 1.02, or 1

Integer for H = 2.18 ÷ 2.18 = 1.00, or 1

Integer for O = 4.443 ÷ 2.18 = $2.0\underline{3}8$, or 2

The empirical formula is thus CHO_2. The formula weight corresponding to this formula may be found by adding the respective atomic weights:

Formula weight = 12.01 amu + 1.008 amu + (2 x 16.00 amu) = 45.02 amu

Dividing the molecular weight by the formula weight gives the number of times the CHO_2 unit occurs in the molecule. Because the molecular weight is 90 amu, this quotient is

90 amu ÷ 45.02 amu = $2.\underline{0}0$, or 2

The molecular formula is thus $(CHO_2)_2$, or $C_2H_2O_4$.

3.70 Assume a sample of 100.0 g of adipic acid. By multiplying this by the percentage composition, we obtain 49.3 g C, 6.9 g H, and 43.8 g O. Convert each of these masses to moles by dividing by the molar mass.

$$\text{Mol C} = 49.3 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.1\underline{0}5 \text{ mol}$$

$$\text{Mol H} = 6.9 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 6.\underline{8}5 \text{ mol}$$

(continued)

$$\text{Mol O} = 43.8 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.7\underline{3}8 \text{ mol (smallest number)}$$

Now, divide each number of moles by the smallest number to obtain the smallest set of integers for the empirical formula.

Tentative integer for C = $4.105 \div 2.738 = 1.4\underline{9}9$, or 1.5

Tentative integer for H = $6.85 \div 2.738 = 2.\underline{5}0$, or 2.5

Tentative integer for O = $2.738 \div 2.738 = 1.0\underline{0}0$, or 1

Because 1.5 and 2.5 are not whole numbers, multiply each tentative integer by two to give the final integers for the empirical formula:

C: $2 \times 1.5 = 3$

H: $2 \times 2.5 = 5$

O: $2 \times 1 = 2$

The empirical formula is thus $C_3H_5O_2$. The formula weight corresponding to this formula may be found by adding the respective atomic weights:

Formula weight = (3 x 12.01 amu) + (5 x 1.008 amu) + (2 x 16.00 amu)

= 73.1 amu

Dividing the molecular weight by the formula weight gives the number of times the CHO_2 unit occurs in the molecule. Because the molecular weight is 146 amu, this quotient is

146 amu $\div$ 73.1 amu = $2.\underline{0}0$, or 2

The molecular formula is thus $(C_3H_5O_2)_2$, or $C_6H_{10}O_4$.

3.71	C_2H_4	+	$3O_2$	$\rightarrow$	$2CO_2$	+	$2H_2O$
	1 molecule C_2H_4	+	3 molecules O_2	$\rightarrow$	2 molecules CO_2	+	2 molecules H_2O
	1 mole C_2H_4	+	3 moles O_2	$\rightarrow$	2 moles CO_2	+	2 moles H_2O
	28.052 g C_2H_4	+	3 x 32.00 g O_2	$\rightarrow$	2 x 44.01 g CO_2	+	2 x 18.016 g H_2O

3.72	$2H_2S$	+	$3O_2$	$\rightarrow$	$2SO_2$	+	$2H_2O$
	2 molecules H_2S	+	3 molecules O_2	$\rightarrow$	2 molecules SO_2	+	2 molecules H_2O
	2 moles H_2S	+	3 moles O_2	$\rightarrow$	2 moles SO_2	+	2 moles H_2O
	2 x 34.06 g H_2S	+	3 x 32.00 g O_2	$\rightarrow$	2 x 64.06 g CO_2	+	2 x 18.016 g H_2O

3.73 By inspecting the balanced equation, obtain a conversion factor of eight mol CO_2 to two mol C_4H_{10}. Multiply the given amount of 0.30 moles of C_4H_{10} by the conversion factor to obtain the moles of H_2O.

$$0.30 \text{ mol } C_4H_{10} \text{ x } \frac{8 \text{ mol } CO_2}{2 \text{ mol } C_4H_{10}} = 1.\underline{2}0 = 1.2 \text{ mol } CO_2$$

3.74 By inspecting the balanced equation, obtain a conversion factor of three mol H_2O to one mol C_2H_5OH. Multiply the given amount of 0.69 mol of C_2H_5OH by the conversion factor to obtain the moles of H_2O.

$$0.69 \text{ mol } C_2H_5OH \text{ x } \frac{3 \text{ mol } H_2O}{1 \text{ mol } C_2H_5OH} = 2.\underline{0}7 = 2.1 \text{ mol } H_2O$$

3.75 By inspecting the balanced equation, obtain a conversion factor of three mol O_2 to two mol Fe_2O_3. Multiply the given amount of 3.91 mol Fe_2O_3 by the conversion factor to obtain moles of O_2.

$$3.91 \text{ mol } Fe_2O_3 \text{ x } \frac{3 \text{ mol } O_2}{2 \text{ mol } Fe_2O_3} = 5.8\underline{6}5 = 5.87 \text{ mol } O_2$$

3.76 By inspecting the balanced equation, obtain a conversion factor of three mol $NiCl_2$ to one mol $Ni_3(PO_4)_2$. Multiply the given amount of 0.479 mol $Ni_3(PO_4)_2$ by the conversion factor to obtain moles of $NiCl_2$.

$$0.479 \text{ mol } Ni_3(PO_4)_2 \text{ x } \frac{3 \text{ mol } NiCl_2}{1 \text{ mol } Ni_3(PO_4)_2} = 1.4\underline{3}7 = 1.44 \text{ mol } NiCl_2$$

3.77 $3\ NO_2 + H_2O \rightarrow 2\ HNO_3 + NO$

three mol of NO_2 are equivalent to two mol of HNO_3 (from equation).

one mol of NO_2 is equivalent to 46.01 g NO_2 (from molecular weight of NO_2).

one mol of HNO_3 is equivalent to 63.02 g HNO_3 (from molecular weight of HNO_3).

$$7.50\ g\ HNO_3 \times \frac{1\ mol\ HNO_3}{63.02\ g\ HNO_3} \times \frac{3\ mol\ NO_2}{2\ mol\ HNO_3} \times \frac{46.01\ g\ NO_2}{1\ mol\ NO_2} = 8.2\underline{1}3$$

$$= 8.21\ g\ NO_2$$

3.78 $2Ca_3(PO_4)_2 + 6SiO_2 + 10C \rightarrow P_4 + 6CaSiO_3 + 10CO$

two mol of $Ca_3(PO_4)_2$ are equivalent to one mol of P_4 (from equation).

one mol of P_4 is equivalent to 123.9 g P_4 (from molecular weight of P_4).

one mol of $Ca_3(PO_4)_2$ is equivalent to 310.2 g $Ca_3(PO_4)_2$ [from molecular weight of $Ca_3(PO_4)_2$].

$$15.0\ g\ P_4 \times \frac{1\ mol\ P_4}{123.88\ g\ P_4} \times \frac{2\ mol\ Ca_3(PO_4)_2}{1\ mol\ P_4} \times \frac{310.18\ g\ Ca_3(PO_4)_2}{1\ mol\ Ca_3(PO_4)_2}$$

$$= 75.\underline{1}1 = 75.1\ g\ Ca_3(PO_4)_2$$

3.79 $WO_3 + 3H_2 \rightarrow W + 3H_2O$

one mol of W is equivalent to three moles of H_2 (from equation).

one mol of H_2 is equivalent to 2.016 g H_2 (from molecular weight of H_2).

one mol of W is equivalent to 183.8 g W (from atomic weight of W).

4.81 kg of H_2 is equivalent to 4.81×10^3 g of H_2.

$$4.81 \times 10^3\ g\ H_2 \times \frac{1\ mol\ H_2}{2.016\ g\ H_2} \times \frac{1\ mol\ W}{3\ mol\ H_2} \times \frac{183.85\ g\ W}{1\ mol\ W} = 1.4\underline{6}2 \times 10^5$$

$$= 1.46 \times 10^5\ g\ W$$

3.80 $4C_3H_6 + 6NO \rightarrow 4C_3H_3N + 6H_2O + N_2$

four mol of C_3H_6 are equivalent to four mol of C_3H_3N (from equation).

one mol of C_3H_6 is equivalent to 42.08 g C_3H_6 (from molecular weight of C_3H_6).

one mol of C_3H_3N is equivalent to 53.06 g C_3H_3N (from molecular weight of C_3H_3N).

651 kg of C_3H_6 are equivalent to 6.51×10^5 g C_3H_6.

$$6.51 \times 10^5 \text{ g } C_3H_6 \times \frac{1 \text{ mol } C_3H_6}{42.08 \text{ g } C_3H_6} \times \frac{4 \text{ mol } C_3H_3N}{4 \text{ mol } C_3H_6} \times \frac{53.06 \text{ g } C_3H_3N}{1 \text{ mol } C_3H_3N}$$

$$= 8.2\underline{0}8 \times 10^5 = 8.21 \times 10^5 \text{ g } C_3H_3N$$

3.81 Write the equation, and set up the calculation below the equation (after calculating the two molecular weights):

$CS_2 + 3Cl_2 \rightarrow CCl_4 + S_2Cl_2$

$$62.7 \text{ g } Cl_2 \times \frac{1 \text{ mol } Cl_2}{70.90 \text{ g } Cl_2} \times \frac{1 \text{ mol } CS_2}{3 \text{ mol } Cl_2} \times \frac{76.15 \text{ g } CS_2}{1 \text{ mol } CS_2} = 22.\underline{4}48$$

$$= 22.4 \text{ g } CS_2$$

3.82 From the molecular models, the balanced chemical equation is

$$4NH_3 + 5O_2 \xrightarrow{Pt} 4NO + 6H_2O$$

The molar mass of NH_3 is 17.03 g/mol, and for O_2, it is 32.00 g/mol. This gives

$$6.1 \text{ g } NH_3 \times \frac{1 \text{ mol } NH_3}{17.03 \text{ g } NH_3} \times \frac{5 \text{ mol } O_2}{4 \text{ mol } NH_3} \times \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2}$$

$$= 14.\underline{3}2 = 14.3 \text{ g } O_2$$

3.83 Write the equation, and set up the calculation below the equation (after calculating the two molecular weights):

$2N_2O_5 \rightarrow 4NO_2 + O_2$

$$1.315 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} \times \frac{4 \text{ mol } NO_2}{1 \text{ mol } O_2} \times \frac{46.01 \text{ g } NO_2}{1 \text{ mol } NO_2}$$

$$= 7.56\underline{2}8 = 7.563 \text{ g } NO_2$$

3.84 Write the equation, and set up the calculation below the equation (after calculating the two formula weights):

$$3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$$

$$5.92\ g\ Cu(NO_3)_2 \times \frac{1\ mol\ Cu(NO_3)_2}{187.56\ g\ Cu(NO_3)_2} \times \frac{2\ mol\ NO}{3\ mol\ Cu(NO_3)_2} \times \frac{30.01\ g\ NO}{1\ mol\ NO}$$

$$= 0.63\underline{1}4 = 0.631\ g\ NO$$

3.85 First determine whether KO_2 or H_2O is the limiting reactant by calculating the moles of O_2 that each would form if it were the limiting reactant. Identify the limiting reactant by the smaller number of moles of O_2 formed.

$$0.15\ mol\ H_2O \times \frac{3\ mol\ O_2}{2\ mol\ H_2O} = 0.2\underline{2}5\ mol\ O_2$$

$$0.25\ mol\ KO_2 \times \frac{3\ mol\ O_2}{4\ mol\ KO_2} = 0.1\underline{8}7\ mol\ O_2\ (KO_2\ \text{is the limiting reactant})$$

The moles of O_2 produced = 0.19 mol.

3.86 First, determine whether NaOH or Cl_2 is the limiting reactant by calculating the moles of NaClO that each would form if it were the limiting reactant. Identify the limiting reactant by the smaller number of moles of NaClO formed.

$$1.23\ mol\ NaOH \times \frac{1\ mol\ NaClO}{2\ mol\ NaOH} = 0.615\ mol\ NaClO\ (\text{smaller number})$$

$$1.47\ mol\ Cl_2 \times \frac{1\ mol\ NaClO}{1\ mol\ Cl_2} = 1.47\ mol\ NaClO$$

NaOH is the limiting reactant, and 0.615 moles of NaClO will form.

3.87 First determine whether CO or H_2 is the limiting reactant by calculating the moles of CH_3OH that each would form if it were the limiting reactant. Identify the limiting reactant by the smaller number of moles of CH_3OH formed. Use the molar mass of CH_3OH to calculate the mass of CH_3OH formed. Then, calculate the mass of the unconsumed reactant.

$$CO + 2H_2 \rightarrow CH_3OH$$

(continued)

$$10.2\text{ g }H_2 \times \frac{1\text{ mol }H_2}{2.016\text{ g }H_2} \times \frac{1\text{ mol }CH_3OH}{2\text{ mol }H_2} = 2.5\underline{2}9\text{ mol }CH_3OH$$

$$35.4\text{ g CO} \times \frac{1\text{ mol CO}}{28.01\text{ g CO}} \times \frac{1\text{ mol }CH_3OH}{1\text{ mol CO}} = 1.2\underline{6}3\text{ mol }CH_3OH$$

CO is the limiting reactant.

$$\text{Mass }CH_3OH\text{ formed} = 1.2\underline{6}3\text{ mol }CH_3OH \times \frac{32.042\text{ g }CH_3OH}{1\text{ mol }CH_3OH}$$

$$= 40.\underline{4}7 = 40.5\text{g }CH_3OH$$

Hydrogen is left unconsumed at the end of the reaction. The mass of H_2 that reacts can be calculated from the moles of product obtained:

$$1.2\underline{6}3\text{ mol }CH_3OH \times \frac{2\text{ mol }H_2}{1\text{mol }CH_3OH} \times \frac{2.016\text{ g }H_2}{1\text{mol }H_2} = 5.0\underline{9}2\text{ g }H_2$$

The unreacted H_2 = 10.2 g total H_2 - 5.092 g reacted H_2 = $5.\underline{1}08$ = 5.1 g H_2.

3.88 First, determine whether CS_2 or O_2 is the limiting reactant by calculating the moles of SO_2 that each would form if it were the limiting reactant. Identify the limiting reactant by the smaller number of moles of SO_2 formed. Use the molar mass of SO_2 to calculate the mass of SO_2 formed. Then, calculate the mass of the unconsumed reactant.

$$CS_2 + 3O_2 \rightarrow CO_2 + 2SO_2$$

$$30.0\text{ g }O_2 \times \frac{1\text{ mol }O_2}{32.00\text{ g }O_2} \times \frac{2\text{ mol }SO_2}{3\text{ mol }O_2} = 0.62\underline{5}0\text{ mol }SO_2$$

$$35.0\text{ g }CS_2 \times \frac{1\text{ mol }CS_2}{76.15\text{ g }CS_2} \times \frac{2\text{ mol }SO_2}{1\text{ mol }CS_2} = 0.91\underline{9}2\text{ mol }SO_2$$

O_2 is the limiting reactant.

$$\text{Mass }SO_2\text{ formed} = 0.62\underline{5}0\text{ mol }SO_2 \times \frac{64.07\text{ g }SO_2}{1\text{mol }SO_2} = 40.\underline{0}4 = 40.0\text{ g}$$

(continued)

CS_2 is left unconsumed at the end of the reaction. The mass of CS_2 that reacts can be calculated from the moles of product obtained:

$$0.62\underline{5}0 \text{ mol } SO_2 \times \frac{1 \text{ mol } CS_2}{2 \text{ mol } SO_2} \times \frac{76.15 \text{ g } CS_2}{1 \text{ mol } CS_2} = 23.\underline{7}9 \text{ g } CS_2$$

The unreacted CS_2 = 30.0 g total CS_2 - 23.$\underline{7}$9 g reacted CS_2 = 6.$\underline{2}$0 = 6.2 g CS_2

3.89 First, determine which of the three reactants is the limiting reactant by calculating the moles of $TiCl_4$ that each would form if it were the limiting reactant. Identify the limiting reactant by the smallest number of moles of $TiCl_4$ formed. Use the molar mass of $TiCl_4$ to calculate the mass of $TiCl_4$ formed.

$$3TiO_2 + 4C + 6Cl_2 \rightarrow 3TiCl_4 + 2CO_2 + 2CO$$

$$4.15 \text{ g } TiO_2 \times \frac{1 \text{ mol } TiO_2}{79.88 \text{ g } TiO_2} \times \frac{3 \text{ mol } TiCl_4}{3 \text{ mol } TiO_2} = 0.051\underline{9}5 \text{ mol } TiCl_4$$

$$5.67 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} \times \frac{3 \text{ mol } TiCl_4}{4 \text{ mol C}} = 0.35\underline{4}07 \text{ mol } TiCl_4$$

$$6.78 \text{ g } Cl_2 \times \frac{1 \text{ mol } Cl_2}{70.90 \text{ g } Cl_2} \times \frac{3 \text{ mol } TiCl_4}{6 \text{ mol } Cl_2} = 0.047\underline{8}1 \text{ mol } TiCl_4$$

Cl_2 is the limiting reactant.

$$\text{Mass } TiCl_4 \text{ formed} = 0.047\underline{8}1 \text{ mol } TiCl_4 \times \frac{189.68 \text{ g } TiCl_4}{1 \text{ mol } TiCl_4} = 9.0\underline{6}8$$

$$= 9.07 \text{ g } TiCl_4$$

3.90 First, determine which of the three reactants is the limiting reactant by calculating the moles of HCN that each would form if it were the limiting reactant. Identify the limiting reactant by the smallest number of moles of HCN formed. Use the molar mass of HCN to calculate the mass of HCN formed.

$$2NH_3 + 3O_2 + 2CH_4 \rightarrow 2HCN + 6H_2O$$

$$11.5 \text{ g } NH_3 \times \frac{1 \text{ mol } NH_3}{17.03 \text{ g } NH_3} \times \frac{2 \text{ mol HCN}}{2 \text{ mol } NH_3} = 0.67\underline{5} \text{ mol HCN}$$

$$10.5 \text{ g } CH_4 \times \frac{1 \text{ mol } CH_4}{16.04 \text{ g } CH_4} \times \frac{2 \text{ mol HCN}}{2 \text{ mol } CH_4} = 0.65\underline{4} \text{ mol HCN}$$

(continued)

$$12.0 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} \times \frac{2 \text{ mol HCN}}{3 \text{ mol } O_2} = 0.25\underline{0}0 \text{ mol HCN}$$

O_2 is the limiting reactant.

$$\text{Mass HCN formed} = 0.2500 \text{ mol HCN} \times \frac{27.03 \text{ g HCN}}{1 \text{ mol HCN}} = 6.5\underline{7}5 = 6.76 \text{ g HCN}$$

3.91 First, determine which of the two reactants is the limiting reactant by calculating the moles of aspirin that each would form if it were the limiting reactant. Identify the limiting reactant by the smallest number of moles of aspirin formed. Use the molar mass of aspirin to calculate the theoretical yield in grams of aspirin. Then calculate the percentage yield.

$$C_7H_6O_3 + C_4H_6O_3 \rightarrow C_9H_8O_4 + C_2H_4O_2$$

$$4.00 \text{ g } C_4H_6O_3 \times \frac{1 \text{ mol } C_4H_6O_3}{102.09 \text{ g } C_4H_6O_3} \times \frac{1 \text{ mol } C_9H_8O_4}{1 \text{ mol } C_4H_6O_3} = 0.039\underline{1}8 \text{ mol } C_9H_8O_4$$

$$2.00 \text{ g } C_7H_6O_3 \times \frac{1 \text{ mol } C_7H_6O_3}{138.12 \text{ g } C_7H_6O_3} \times \frac{1 \text{ mol } C_9H_8O_4}{1 \text{ mol } C_7H_6O_3} = 0.014\underline{4}8 \text{ mol } C_9H_8O_4$$

Thus, $C_7H_6O_3$ is the limiting reactant. The theoretical yield of $C_9H_8O_4$ is

$$0.014\underline{4}8 \text{ mol } C_9H_8O_4 \times \frac{180.15 \text{ g } C_9H_8O_4}{1 \text{ mol } C_9H_8O_4} = 2.6\underline{0}9 \text{ g } C_9H_8O_4$$

The percentage yield is

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{1.86 \text{ g}}{2.609 \text{ g}} \times 100\% = 71.\underline{2}9 = 71.3\%$$

3.92 First, determine which of the two reactants is the limiting reactant by calculating the moles of methyl salicylate that each would form if it were the limiting reactant. Identify the limiting reactant by the smallest number of moles of methyl salicylate formed. Use the molar mass of methyl salicylate to calculate the theoretical yield in grams of methyl salicylate. Then calculate the percentage yield.

$$C_7H_6O_3 + CH_3OH \rightarrow C_8H_8O_3 + H_2O$$

$$11.20 \text{ g } CH_3OH \times \frac{1 \text{ mol } CH_3OH}{32.04 \text{ g } CH_3OH} \times \frac{1 \text{ mol } C_8H_8O_3}{1 \text{ mol } CH_3OH} = 0.34\underline{9}6 \text{ mol } C_8H_8O_3$$

(continued)

$$1.50 \text{ g } C_7H_6O_3 \times \frac{1 \text{ mol } C_7H_6O_3}{138.12 \text{ g } C_7H_6O_3} \times \frac{1 \text{ mol } C_8H_8O_3}{1 \text{ mol } C_7H_6O_3}$$

$$= 0.010\underline{8}6 \text{ mol } C_8H_8O_3$$

Thus, $C_7H_6O_3$ is the limiting reactant. The theoretical yield of $C_8H_8O_3$ is

$$0.010\underline{8}6 \text{ mol } C_8H_8O_3 \times \frac{152.14 \text{ g } C_8H_8O_3}{1 \text{ mol } C_8H_8O_3} = 1.6\underline{5}2 \text{ g } C_8H_8O_3$$

The percentage yield is

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{1.27 \text{ g}}{1.651 \text{ g}} \times 100\% = 76.\underline{9}2 = 76.9\%$$

■ Solutions to General Problems

3.93 For 1 mol of caffeine, there are eight mol of C, ten mol of H, four mol of N, and two mol of O. Convert these amounts to masses by multiplying them by their respective molar masses:

8 mol C x 12.01 g C/1 mol C	=	96.08 g C
10 mol H x 1.008 g H/1 mol H	=	10.08 g H
4 mol N x 14.01 g N/1 mol N	=	56.04 g N
2 mol O x 16.00 g O/1 mol O	=	32.00 g O
1 mol of caffeine (total)	=	194.20 g (molar mass)

Each mass percentage is calculated by dividing the mass of the element by the molar mass of caffeine and multiplying by 100 percent: Mass percentage = (mass element ÷ mass caffeine) x 100%.

Mass percentage C = (96.08 g ÷ 194.20 g) x 100% = 49.5% (3 s.f.)

Mass percentage H = (10.08 g ÷ 194.20 g) x 100% = 5.19% (3 s.f.)

Mass percentage N = (56.04 g ÷ 194.20 g) x 100% = 28.9% (3 s.f.)

Mass percentage O = (32.00 g ÷ 194.20 g) x 100% = 16.5% (3 s.f.)

3.94 For one mol of morphine, there are seventeen mol of C, nineteen mol of H, one mol of N, and three mol of O. Convert these amounts to masses by multiplying by the respective molar masses:

17 mol C x 12.01 g C/1 mol C	=	204.17 g C
19 mol H x 1.008 g H/1 mol H	=	19.15 g H
1 mol N x 14.01 g N/1 mol N	=	14.01 g N
3 mol O x 16.00 g O/1 mol O	=	<u>48.00 g O</u>
1 mol of morphine (total)	=	285.33 g (molar mass)

Each mass percentage is calculated by dividing the mass of the element by the molar mass of morphine and multiplying by 100 percent: Mass percentage = (mass element ÷ mass morphine) x 100%.

Mass percentage C = (204.17 g ÷ 285.33 g) x 100% = 71.6% (3 s.f.)

Mass percentage H = (19.15 g ÷ 285.33 g) x 100% = 6.71% (3 s.f.)

Mass percentage N = (14.01 g ÷ 285.33 g) x 100% = 4.91% (3 s.f.)

Mass percentage O = (48.00 g ÷ 285.33 g) x 100% = 16.8% (3 s.f.)

3.95 Assume a sample of 100.0 g of dichlorobenzene. By multiplying this by the percentage composition, we obtain 49.1 g C, 2.7 g of H, and 48.2 g of Cl. Convert each mass to moles by dividing by the molar mass:

$$49.1 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.0\underline{8}8 \text{ mol C}$$

$$2.7 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 2.\underline{6}8 \text{ mol H}$$

$$48.2 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 1.3\underline{6}0 \text{ mol Cl}$$

Divide each number of moles by the smallest number to obtain the smallest set of integers for the empirical formula.

Integer for C = 4.088 mol ÷ 1.360 mol = 3.00, or 3

Integer for H = 2.68 mol ÷ 1.360 mol = 1.97, or 2

Integer for Cl = 1.360 mol ÷ 1.360 mol = 1.00, or 1

(continued)

The empirical formula is thus C_3H_2Cl. Find the formula weight by adding the atomic weights:

$$\text{Formula weight} = (3 \times 12.01 \text{ amu}) + (2 \times 1.008 \text{ amu}) + 35.45 \text{ amu}$$

$$= 73.4\underline{9}6 = 73.50 \text{ amu}$$

Divide the molecular weight by the formula weight to find the number of times the C_3H_2Cl unit occurs in the molecule. Because the molecular weight is 147 amu, this quotient is

$$147 \text{ amu} \div 73.50 \text{ amu} = 2.00, \text{ or } 2$$

The molecular formula is $(C_3H_2Cl)_2$, or $C_6H_4Cl_2$.

3.96 Assume a sample of 100.0 g of sorbic acid. By multiplying this by the percentage composition, we obtain 64.3 g C, 7.2 g H, and 28.5 g O. Convert each mass to moles by dividing by the molar mass.

$$64.3 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 5.3\underline{5}3 \text{ mol C}$$

$$7.2 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 7.\underline{1}4 \text{ mol H}$$

$$28.5 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.7\underline{8}1 \text{ mol O}$$

Divide each number of moles by the smallest number to obtain the smallest set of integers for the empirical formula.

$$\text{Integer for C} = 5.353 \text{ mol} \div 1.781 \text{ mol} = 3.01, \text{ or } 3$$

$$\text{Integer for H} = 7.14 \text{ mol} \div 1.781 \text{ mol} = 4.0, \text{ or } 4$$

$$\text{Integer for O} = 1.781 \text{ mol} \div 1.781 \text{ mol} = 1.00, \text{ or } 1$$

The empirical formula is thus C_3H_4O.

Find the formula weight by adding the atomic weights:

$$\text{Formula weight} = (3 \times 12.01 \text{ amu}) + (4 \times 1.008 \text{ amu}) + 16.00 \text{ amu}$$

$$= 56.0\underline{6}2 = 56.06 \text{ amu}$$

(continued)

Divide the molecular weight by the formula weight to find the number of times the C_3H_4O unit occurs in the molecule. Because the molecular weight was given as 112 amu, this quotient is

$$112 \text{ amu} \div 56.06 \text{ amu} = 2.00, \text{ or } 2$$

The molecular formula is $(C_3H_4O)_2$, or $C_6H_8O_2$.

3.97 Find the percent composition of C and S from the analysis:

$$0.01665 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2} \times \frac{1 \text{ mol C}}{1 \text{ mol } CO_2} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 0.00454\underline{4} \text{ g C}$$

$$\text{Percent C} = (0.004544 \text{ g C} \div 0.00796 \text{ g comp.}) \times 100\% = 57.\underline{0}9\%$$

$$0.01196 \text{ g } BaSO_4 \times \frac{1 \text{ mol } BaSO_4}{233.39 \text{ g } BaSO_4} \times \frac{1 \text{ mol S}}{1 \text{ mol } BaSO_4} \times \frac{32.07 \text{ g S}}{1 \text{ mol S}}$$

$$= 0.0016\underline{4}3 \text{ g S}$$

$$\text{Percent S} = (0.001643 \text{ g S} \div 0.00431 \text{ g comp.}) \times 100\% = 38.\underline{1}2\%$$

$$\text{Percent H} = 100.00\% - (57.09 + 38.12)\% = 4.\underline{7}9\%$$

We now obtain the empirical formula by calculating moles from the grams corresponding to each mass percentage of element:

$$57.09 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.75\underline{4} \text{ mol C}$$

$$38.12 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g C}} = 1.18\underline{9} \text{ mol S}$$

$$4.79 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.7\underline{5}2 \text{ mol H}$$

Dividing the moles of the elements by the smallest number (1.189), we obtain for C: 3.997, or 4; for S: 1.000, or 1; and for H: 3.996, or 4. Thus, the empirical formula is C_4H_4S (formula weight = 84). Because the formula weight was given as 84 amu, the molecular formula is also C_4H_4S.

3.98 Find the percentage composition of H and N from the analysis:

$$0.00663 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2O} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.00074\underline{1}7 \text{ g H}$$

1.46 mg N_2 from the analysis is equivalent to 0.00146 g N.

$$\text{Percent H} = (0.0007417 \text{ g H} \div 0.00971 \text{ g comp.}) \times 100\% = 7.6\underline{4}\%$$

$$\text{Percent N} = (0.00146 \text{ g N} \div 0.00971 \text{ g comp.}) \times 100\% = 15.\underline{0}\%$$

$$\text{Percent C} = 100.00\% - (7.64 + 15.0)\% = 77.\underline{4}\%$$

Calculate the moles from grams to obtain the empirical formula:

$$77.4 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 6.\underline{4}4 \text{ mol C}$$

$$7.64 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 7.5\underline{8} \text{ mol H}$$

$$15.0 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 1.0\underline{7} \text{ mol N}$$

Dividing the moles of elements by the smallest number (1.07) gives for C: 6.02, or 6; for H: 7.08, or 7; and for N: 1.00, or 1. The empirical formula is thus C_6H_7N (formula weight = 93 amu). Because the molecular weight was given as 93 amu, the molecular formula is also C_6H_7N.

3.99 For g $CaCO_3$, use this equation: $CaCO_3 + H_2C_2O_4 \rightarrow CaC_2O_4 + H_2O + CO_2$.

$$0.472 \text{ g } CaC_2O_4 \times \frac{1 \text{ mol } CaC_2O_4}{128.10 \text{ g } CaC_2O_4} \times \frac{1 \text{ mol } CaCO_3}{1 \text{ mol } CaC_2O_4} \times \frac{100.09 \text{ g } CaCO_3}{1 \text{ mol } CaCO_3}$$

$$= 0.36\underline{8}8 \text{ g } CaCO_3$$

$$\text{Mass percentage } CaCO_3 = \frac{\text{mass } CaCO_3}{\text{mass limestone}} \times 100\% = \frac{0.36\underline{8}8 \text{ g}}{0.438 \text{ g}} \times 100\%$$

$$= 84.\underline{1}9 = 84.2\%$$

3.100 For the mass of TiO_2, use this equation: $TiO_2 + C + 2Cl_2 \rightarrow TiCl_4 + CO_2$.

$$35.4\text{ g }TiCl_4 \times \frac{1\text{ mol }TiCl_4}{189.68\text{ g }TiCl_4} \times \frac{1\text{ mol }TiO_2}{1\text{ mol }TiCl_4} \times \frac{79.88\text{ g }TiO_2}{1\text{ mol }TiO_2} = 14.\underline{9}1\text{ g }TiO_2$$

$$\text{Mass percent }TiO_2 = \frac{\text{mass }TiO_2}{\text{mass rutile}} \times 100\% = \frac{14.\underline{9}1\text{ g}}{17.4\text{ g}} \times 100\%$$

$$= 85.\underline{6}8 = 85.7\%$$

3.101 Calculate the theoretical yield using this equation: $2C_2H_4 + O_2 \rightarrow 2C_2H_4O$.

$$10.6\text{ g }C_2H_4 \times \frac{1\text{ mol }C_2H_4}{28.05\text{ g }C_2H_4} \times \frac{2\text{ mol }C_2H_4O}{2\text{ mol }C_2H_4} \times \frac{44.05\text{ g }C_2H_4O}{1\text{ mol }C_2H_4O}$$

$$= 16.\underline{6}5\text{ g }C_2H_4O$$

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{9.91\text{ g}}{16.65\text{ g}} \times 100\% = 59.\underline{5}3 = 59.5\%$$

3.102 Calculate the theoretical yield using this equation:

$$C_6H_6 + HNO_3 \rightarrow C_6H_5NO_2 + H_2O$$

$$22.4\text{ g }C_6H_6 \times \frac{1\text{ mol }C_6H_6}{78.11\text{ g }C_6H_6} \times \frac{1\text{ mol }C_6H_5NO_2}{1\text{ mol }C_6H_6} \times \frac{123.11\text{ g }C_6H_5NO_2}{1\text{ mol }C_6H_5NO_2}$$

$$= 35.\underline{3}0\text{ g }C_6H_5NO_2$$

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{31.6\text{ g}}{35.\underline{3}0\text{ g}} \times 100\% = 89.\underline{5}1 = 89.5\%$$

3.103 To find Zn, use these equations:

$$2C + O_2 \rightarrow 2CO \text{ and } ZnO + CO \rightarrow Zn + CO_2$$

Two mol C produces two mol CO; because one mol ZnO reacts with one mol CO, two mol ZnO will react with two mol CO. Thus, two mol C is equivalent to two mol ZnO, or one mol C is equivalent to one mol ZnO.

(continued)

Using this to calculate mass of C from mass of ZnO, we have

$$75.0\text{ g ZnO} \times \frac{1\text{ mol ZnO}}{81.39\text{ g ZnO}} \times \frac{1\text{ mol C}}{1\text{ mol ZnO}} \times \frac{12.01\text{ g C}}{1\text{ mol C}} = 11.\underline{0}7\text{ g C}$$

Thus, all of the ZnO is used up in reacting with just 11.07 g of C, making ZnO the limiting reactant. Use the mass of ZnO to calculate the mass of Zn formed:

$$75.0\text{ g ZnO} \times \frac{1\text{ mol ZnO}}{81.39\text{ g ZnO}} \times \frac{1\text{ mol Zn}}{1\text{ mol ZnO}} \times \frac{65.39\text{ g Zn}}{1\text{ mol Zn}} = 60.\underline{2}56 = 60.3\text{ g Zn}$$

3.104 To find CH_4, use these equations:

$$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$$

$$2NO + 2CH_4 \rightarrow 2HCN + 2H_2O + H_2$$

Four mol NH_3 produces four mol NO; because two mol CH_4 reacts with two mol NO, four mol CH_4 will react with four mol NO. Thus, four mol NH_3 is equivalent to four mol CH_4. Using this to calculate the mass of CH_4 from the mass of NH_3, we have

$$24.2\text{ g NH}_3 \times \frac{1\text{ mol NH}_3}{17.03\text{ g NH}_3} \times \frac{4\text{ mol CH}_4}{4\text{ mol NH}_3} \times \frac{16.04\text{ g CH}_4}{1\text{ mol CH}_4} = 22.\underline{8}\text{ g CH}_4$$

Thus, all of the NH_3 is used up in reacting with just 22.8 g of CH_4, making NH_3 the limiting reactant. Use the mass of NH_3 to calculate the mass of HCN formed:

$$24.2\text{ g NH}_3 \times \frac{1\text{ mol NH}_3}{17.03\text{ g NH}_3} \times \frac{2\text{ mol CH}_4}{2\text{ mol NH}_3} \times \frac{27.03\text{ g HCN}}{1\text{ mol HCN}}$$

$$= 38.\underline{4}1 = 38.4\text{ g HCN}$$

3.105 For $CaO + 3C \rightarrow CaC_2 + CO$, find the limiting reactant in terms of moles of CaC_2 obtainable:

$$\text{Mol CaC}_2 = 2.60 \times 10^3\text{ g C} \times \frac{1\text{ mol C}}{12.01\text{ g C}} \times \frac{1\text{ mol CaC}_2}{3\text{ mol C}} = 72.\underline{1}6\text{ mol}$$

$$\text{Mol CaC}_2 = 2.60 \times 10^3\text{ g CaO} \times \frac{1\text{ mol CaO}}{56.08\text{ g CaO}} \times \frac{1\text{ mol CaC}_2}{1\text{ mol CaO}} \quad 46.\underline{3}62\text{ mol}$$

(continued)

Because CaO is the limiting reactant, calculate the mass of CaC_2 from it:

$$\text{Mass } CaC_2 = 46.\underline{3}62 \text{ mol } CaC_2 \times \frac{64.10 \text{ g } CaC_2}{1 \text{ mol } CaC_2} = 2.9\underline{7}1 \times 10^3$$

$$= 2.97 \times 10^3 \text{ g } CaC_2$$

3.106 For $CaF_2 + H_2SO_4 \rightarrow 2HF + CaSO_4$, find the limiting reactant in terms of moles of HF obtainable:

$$\text{Mol HF} = 12.8 \text{ g } CaF_2 \times \frac{1 \text{ mol } CaF_2}{78.08 \text{ g } CaF_2} \times \frac{2 \text{ mol HF}}{1 \text{ mol } CaF_2} = 0.32\underline{7}9 \text{ mol}$$

$$\text{Mol HF} = 13.2 \text{ g } H_2SO_4 \times \frac{1 \text{ mol } H_2SO_4}{98.09 \text{ g } H_2SO_4} \times \frac{2 \text{ mol HF}}{1 \text{ mol } H_2SO_4} = 0.26\underline{9}14 \text{ mol}$$

Because H_2SO_4 is the limiting reactant, calculate the mass of HF from it:

$$\text{Mass HF} = 0.26914 \text{ mol HF} \times \frac{20.01 \text{ g HF}}{1 \text{ mol HF}} = 5.3\underline{8}54 = 5.39 \text{ g HF}$$

3.107 From the equation $2Na + H_2O \rightarrow 2NaOH + H_2$, convert the mass of H_2 to mass of Na, and then use the mass to calculate the percentage:

$$0.108 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.016 \text{ g } H_2} \times \frac{2 \text{ mol Na}}{1 \text{ mol } H_2} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 2.4\underline{6}3 \text{ g Na}$$

$$\text{Percent Na} = \frac{\text{mass Na}}{\text{mass amalgam}} \times 100\% = \frac{2.463 \text{ g}}{15.23 \text{ g}} \times 100\% = 16.\underline{1}7 = 16.2\%$$

3.108 From the equation $CaCO_3 \rightarrow CaO + CO_2$, convert the mass of CO_2 to mass of $CaCO_3$, and then use the mass to calculate the percentage:

$$0.00395 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2} \times \frac{1 \text{ mol } CaCO_3}{1 \text{ mol } CO_2} \times \frac{100.09 \text{ g } CaCO_3}{1 \text{ mol } CaCO_3}$$

$$= 0.0089\underline{8}3 \text{ g } CaCO_3$$

(continued)

$$\text{Percent } CaCO_3 = \frac{\text{mass } CaCO_3}{\text{mass sandstone}} \times 100\% = \frac{0.0089\underline{8}3 \text{ g}}{0.0187 \text{ g}} \times 100\%$$

$$= 48.\underline{0}39 = 48.0\%$$

Because the sandstone contains only SiO_2 and $CaCO_3$, the difference between 100 percent and the percentage of $CaCO_3$ is the percentage of SiO_2:

$$\text{Percent } SiO_2\text{(silica)} = 100.00\% - 48.\underline{0}39\% = 51.\underline{9}6 = 52.0\%$$

■ Solutions to Cumulative-Skills Problems

3.109 Let y equal the mass of CuO in the mixture. Then 0.500 g - y equals the mass of Cu_2O in the mixture. Multiplying the appropriate conversion factors for Cu times the mass of each oxide will give one equation in one unknown for the mass of 0.425 g Cu:

$$0.425 = y\left[\frac{63.55 \text{ g Cu}}{79.55 \text{ g CuO}}\right] + (0.500 - y)\left[\frac{127.10 \text{ g Cu}}{143.10 \text{ g } Cu_2O}\right]$$

Simplifying the equation by dividing the conversion factors and combining terms gives:

$$0.425 = 0.798\underline{8}7\ y + 0.8881\underline{9}0\ (0.500 - y)$$

$$0.08932\ y = 0.01\underline{9}095$$

$$y = 0.2\underline{1}39 = 0.21 \text{ g} = \text{mass of CuO}$$

3.110 Let y equal the mass of Fe_2O_3 in the mixture. Then 0.500 g - y equals the mass of FeO in the mixture. The mass of Fe in the mixture is 0.720 x 0.500 g = 0.360 g. Multiplying the appropriate conversion factors for Fe times the mass of each oxide will give one equation in one unknown for the mass of 0.360 g Fe in the mixture:

$$0.360 = y\left[\frac{111.70 \text{ g Fe}}{159.70 \text{ g } Fe_2O_3}\right] + (0.500 - y)\left[\frac{55.85 \text{ g Fe}}{71.85 \text{ g FeO}}\right]$$

Simplifying the equation by dividing the conversion factors and combining terms gives:

$$0.360 = 0.6994\ y + 0.7773\ (0.500 - y)$$

$$0.07790\ y = 0.02865$$

$$y = 0.36\underline{7}7 = 0.368 \text{ g} = \text{mass of } Fe_2O_3$$

3.111 If one heme molecule contains one iron atom, then the number of moles of heme in 35.2 mg heme must be the same as the number of moles of iron in 3.19 mg of iron. Start by calculating the moles of Fe (equals moles heme):

$$3.19 \times 10^{-3}\text{ g Fe} \times \frac{1\text{ mol Fe}}{55.85\text{ g Fe}} = 5.7\underline{1}2 \times 10^{-5}\text{ mol Fe or heme}$$

$$\text{Molar mass of heme} = \frac{35.2 \times 10^{-3}\text{ g}}{5.712 \times 10^{-5}\text{ mol}} = 61\underline{6}.2 = 616\text{ g/mol}$$

The molecular weight of heme is 616 amu.

3.112 Convert the mass of $BaSO_4$ to mass of S to find the percentage of sulfur:

$$0.00546\text{ g BaSO}_4 \times \frac{1\text{ mol BaSO}_4}{233.40\text{ g BaSO}_4} \times \frac{1\text{ mol S}}{1\text{ mol BaSO}_4} \times \frac{32.07\text{ g S}}{1\text{ mol S}}$$

$$= 0.7502 \times 10^{-3}\text{ g S}$$

$$\text{Mass percentage S} = \frac{0.7502 \times 10^{-3}\text{ g S}}{8.19 \times 10^{-3}\text{ g pen. V}} \times 100\% = 9.1\underline{6}0 = 9.16\%$$

Convert the mass of S to moles; then recognizing that moles of S equals moles of pen. V, use that number of moles to calculate the molar mass:

$$0.7502 \times \text{\`{}}0^{-3}\text{ g S} \times \frac{1\text{ mol S}}{32.07\text{ g S}} = 2.3\underline{3}9 \times 10^{-5}\text{ mol S}$$

$$\text{Molar mass pen. V} = \frac{8.19 \times 10^{-3}\text{ g}}{2.339 \times 10^{-5}\text{ mol}} = 35\underline{0}.1 = 350.\text{ g/mol}$$

The molecular weight of penicillin V is 350 amu.

3.113 Use the data to find the molar mass of the metal and anion. Start with X_2.

Mass X_2 in MX = 4.52 g - 3.41 g = 1.11 g

Molar mass X_2 = 1.11 g ÷ 0.0158 mol = $70.\underline{2}5$ g/mol

Molar mass X = 70.25 ÷ 2 = $35.\underline{1}4$ = 35.1 g/mol

(continued)

Thus X is Cl, chlorine.

Moles of M in 4.52 g MX = 0.0158 x 2 = 0.0316 mol

Molar mass of M = 3.41 g $\div$ 0.0316 mol = $10\underline{7}.9$ = 108 g/mol

Thus M is Ag, silver.

3.114 Use the data to find the molar mass of the metal and the anion. Start with M^{2+}.

Mol M^{2+} in MX_2 = 0.158 mol $\div$ 2 = 0.0790 mol

Molar mass of M^{2+} = 1.92 g $\div$ 0.0790 mol = $24.\underline{3}0$ = 24.3 g/mol

Thus M^{2+} is Mg^{2+}.

Now, find the mass of MX_2, and then find the molar mass of X:

(1 - 0.868) mass MX_2 = 1.92 g, or mass MX_2 = $14.\underline{5}45$ g

Mass of X in MX_2 = $14.\underline{5}45$ g - 1.92 g = $12.\underline{6}25$ g

The moles of X in MX_2 = 0.158 mol

Molar mass of X = $12.\underline{6}25$ g $\div$ 0.158 mol = $79.\underline{9}05$ = 79.9 g/mol

Thus X is Br (molar mass 79.90).

3.115 After finding the volume of the alloy, convert it to mass Fe using density and percent Fe. Then use Avogadro's number and the atomic weight for the number of atoms.

Vol. = 10.0 cm x 20.0 cm x 15.0 cm = 3.00×10^3 cm^3

$$\text{Mass Fe} = 3.00 \times 10^3 \text{ cm}^3 \times \frac{8.17 \text{ g alloy}}{1 \text{ cm}^3} \times \frac{54.7 \text{ g Fe}}{100.0 \text{ g alloy}} = 1.3\underline{4}07 \times 10^4 \text{ g}$$

$$\text{No. of Fe atoms} = 1.3\underline{4}07 \times 10^{-4} \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \times \frac{6.022 \times 10^{23} \text{ Fe atoms}}{1 \text{ mol Fe}}$$

$$= 1.4\underline{4}56 \times 10^{26} = 1.45 \times 10^{26} \text{ Fe atoms}$$

3.116 After finding the volume of the cylinder, convert it to mass Co using density and percent Co. Then use Avogadro's number and the atomic weight for the number of atoms.

$$\text{Vol.} = 3.1416 \times (2.50 \text{ cm})^2 \times 10 \text{ cm} = 196.35 \text{ cm}^3$$

$$\text{Mass Co} = 1.9635 \text{ cm}^3 \times \frac{8.20 \text{ g alloy}}{1 \text{ cm}^3} \times \frac{12.0 \text{ g Co}}{100.0 \text{ g alloy}} = 1.9\underline{3}2 \times 10^2 \text{ g}$$

$$\text{No. of Co atoms} = 1.9\underline{3}2 \times 10^2 \text{ g Co} \times \frac{1 \text{ mol Co}}{58.93 \text{ g Co}} \times \frac{6.022 \times 10^{23} \text{ Co atoms}}{1 \text{ mol Co}}$$

$$= 1.9\underline{7}4 \times 10^{24} = 1.97 \times 10^{24} \text{ Co atoms}$$

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