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. \( \text{NO}_2 \)
   
   \[
   \begin{align*}
   1 \times \text{AW of N} &= 14.067 \text{ amu} \\
   2 \times \text{AW of O} &= 2 \times 15.9994 = 31.9988 \text{ amu} \\
   \text{MW of NO}_2 &= 46.0055 = 46.0 \text{ amu (3 s.f.)}
   \end{align*}
   \]

b. \( \text{C}_6\text{H}_12\text{O}_6 \)
   
   \[
   \begin{align*}
   6 \times \text{AW of C} &= 6 \times 12.011 = 72.066 \text{ amu} \\
   12 \times \text{AW of H} &= 12 \times 1.0079 = 12.0948 \text{ amu} \\
   6 \times \text{AW of O} &= 6 \times 15.9994 = 95.9964 \text{ amu} \\
   \text{MW of C}_6\text{H}_{12}\text{O}_6 &= 180.1572 \text{ amu} = 180. \text{ amu (3 s.f.)}
   \end{align*}
   \]

c. \( \text{NaOH} \)
   
   \[
   \begin{align*}
   1 \times \text{AW of Na} &= 22.98977 \text{ amu} \\
   1 \times \text{AW of O} &= 15.9994 \text{ amu} \\
   1 \times \text{AW of H} &= 1.0079 \text{ amu} \\
   \text{MW of NaOH} &= 39.9971 \text{ amu} = 40.0 \text{ amu (3 s.f.)}
   \end{align*}
   \]

d. \( \text{Mg(OH)}_2 \)
   
   \[
   \begin{align*}
   1 \times \text{AW of Mg} &= 24.305 \text{ amu} \\
   2 \times \text{AW of O} &= 2 \times 15.9994 = 31.9988 \text{ amu} \\
   2 \times \text{AW of H} &= 2 \times 1.0079 = 2.0158 \text{ amu} \\
   \text{MW of Mg(OH)}_2 &= 58.3196 \text{ amu} = 58.3 \text{ amu (3 s.f.)}
   \end{align*}
   \]
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$_2$SO$_4$. 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 x 10$^{23}$ Ca atoms.

\[
\text{Mass of one Ca} = \frac{40.08 \text{ g}}{1 \text{ mol Ca}} = \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atoms}} = 6.656 \times 10^{-23}
\]

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

b. The molecular weight of C$_2$H$_5$OH, or C$_2$H$_6$O, = (2 x 12.01) + (6 x 1.008) + 16.00 = 46.068. Its molar mass = 46.07 g/mol, and one mol = 6.022 x 10$^{23}$ molecules of C$_2$H$_6$O.

\[
\text{Mass of one C}_2\text{H}_6\text{O} = \frac{46.07 \text{ g}}{1 \text{ mol C}_2\text{H}_6\text{O}} = \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atoms}} = 7.6503 \times 10^{-23} = 7.650 \times 10^{-23} \text{ g/molecule}
\]

3.4  The molar mass of H$_2$O$_2$ is 34.02 g/mol. Therefore,

\[
0.909 \text{ mol H}_2\text{O}_2 \times \frac{34.02 \text{ g H}_2\text{O}_2}{1 \text{ mol H}_2\text{O}_2} = 30.92 = 30.9 \text{ g H}_2\text{O}_2
\]

3.5  The molar mass of HNO$_3$ is 63.01 g/mol. Therefore,

\[
28.5 \text{ g HNO}_3 \times \frac{1 \text{ mol HNO}_3}{63.01 \text{ g HNO}_3} = 0.4523 = 0.452 \text{ mol HNO}_3
\]

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.

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

\[
= 1.248 \times 10^{21} = 1.2 \times 10^{21} \text{ HCN molecules}
\]
3.7 The molecular weight of \( \text{NH}_4\text{NO}_3 = 80.05 \); thus, its molar mass = 80.05 g/mol. Hence

\[
\text{Percent N} = \frac{28.02 \text{ g}}{80.05 \text{ g}} \times 100\% = 35.00 \% = 35.0\%
\]

\[
\text{Percent H} = \frac{4.032 \text{ g}}{80.05 \text{ g}} \times 100\% = 5.036 \% = 5.04\%
\]

\[
\text{Percent O} = \frac{48.00 \text{ g}}{80.05 \text{ g}} \times 100\% = 59.96 \% = 60.0\%
\]

3.8 From the previous exercise, \( \text{NH}_4\text{NO}_3 \) is 35.0 percent N (fraction N = 0.350), so the mass of N in 48.5 g of \( \text{NH}_4\text{NO}_3 \) is

\[
48.5 \text{ g} \text{NH}_4\text{NO}_3 \times (0.350 \text{ g N} / 1 \text{ g NH}_4\text{NO}_3) = 16.975 = 17.0 \text{ g N}
\]

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:

\[
5.80 \times 10^{-3} \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 1.583 \times 10^{-3} \text{ g C}
\]

Do the same series of calculations for water, noting that one mol H\(_2\)O contains two mol H.

\[
1.58 \times 10^{-3} \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 1.767 \times 10^{-4} \text{ g H}
\]

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

\[
\text{Percent C} = \frac{1.583 \text{ mg}}{3.87 \text{ mg}} \times 100\% = 40.90 \% = 40.9\% \text{ C}
\]

\[
\text{Percent H} = \frac{0.1767 \text{ mg}}{3.87 \text{ mg}} \times 100\% = 4.5658 \% = 4.57\% \text{ H}
\]

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

\[
\text{Percent O} = 100.000\% - (40.90 + 4.5658) = 54.5342 = 54.5\% \text{ O}
\]
3.10 Convert the masses to moles that are proportional to the subscripts in the empirical formula:

\[
33.4 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = 1.0414 \text{ mol S}
\]

\[
(83.5 - 33.4) \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.132 \text{ mol O}
\]

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

For O: \[\frac{3.1312}{1.0414} = 3.01\]
For S: \[\frac{1.0414}{1.0414} = 1.00\]

The empirical formula is \(\text{SO}_3\).

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:

\[
68.8 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 5.729 \text{ mol C}
\]

\[
5.0 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.96 \text{ mol H}
\]

\[
26.2 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.638 \text{ mol O}
\]

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:

For C: \[\frac{5.729}{1.638} = 3.497\]
For H: \[\frac{4.96}{1.638} = 3.03\]

For O: \[\frac{1.638}{1.638} = 1.000\]

Rounding off, we obtain \(\text{C}_{3.5}\text{H}_{3.0}\text{O}_{1.0}\). Multiplying the numbers by two gives whole numbers for an empirical formula of \(\text{C}_7\text{H}_6\text{O}_2\).
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:

\[
\begin{align*}
54.5 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} &= 4.537 \text{ mol C} \\
9.2 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} &= 9.12 \text{ mol H} \\
36.3 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} &= 2.268 \text{ mol O}
\end{align*}
\]

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:

For C: \(\frac{4.537}{2.268} = 2.000\)

For H: \(\frac{9.12}{2.268} = 4.02\)

For O: \(\frac{2.268}{2.268} = 1.000\)

Rounding off, we obtain \(\text{C}_2\text{H}_4\text{O}\), the empirical formula, which is also the molecular formula.

3.13 \(\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}\)

1 molec. (mol) \(\text{H}_2 \) + 1 molec. (mol) \(\text{Cl}_2 \) \(\rightarrow\) 2 molec. (mol) \(\text{HCl}\) (molec., mole interp.)

2.016 g \(\text{H}_2 \) + 70.9 g \(\text{Cl}_2 \) \(\rightarrow\) 2 x 36.5 g \(\text{HCl}\) (mass interp.)

3.14 Equation: \(\text{Na} + \text{H}_2\text{O} \rightarrow \frac{1}{2}\text{H}_2 + \text{NaOH}\), or \(2\text{Na} + 2\text{H}_2\text{O} \rightarrow \text{H}_2 + 2\text{NaOH}\).

From this equation, one mol of Na corresponds to one-half mol of \(\text{H}_2\), or two mol of Na corresponds to one mol of \(\text{H}_2\). Therefore,

\[
7.81 \text{ g } \text{H}_2 \times \frac{1 \text{ mol } \text{H}_2}{2.016 \text{ g } \text{H}_2} \times \frac{2 \text{ mol Na}}{1 \text{ mol } \text{H}_2} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 178.1 = 178 \text{ g Na}
\]

3.15 Balanced equation: \(2\text{ZnS} + 3\text{O}_2 \rightarrow 2\text{ZnO} + 2\text{SO}_2\)

Convert grams of ZnS to moles of ZnS. Then, determine the relationship between ZnS and \(\text{O}_2\) (2ZnS is equivalent to 3O\(_2\)). Finally, convert to mass \(\text{O}_2\).

\[
5.00 \times 10^3 \text{ g ZnS} \times \frac{1 \text{ mol ZnS}}{97.46 \text{ g ZnS}} \times \frac{3 \text{ mol } \text{O}_2}{2 \text{ mol ZnS}} \times \frac{32.00 \text{ g } \text{O}_2}{1 \text{ mol } \text{O}_2} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 2.463 = 2.46 \text{ kg } \text{O}_2
\]
3.16 Balanced equation: $2\text{HgO} \rightarrow 2\text{Hg} + \text{O}_2$

Convert the mass of $\text{O}_2$ to mol of $\text{O}_2$. Using the fact that one mol of $\text{O}_2$ is equivalent to two mol of $\text{Hg}$, determine the number of mol of $\text{Hg}$, and convert to mass of $\text{Hg}$.

$$
6.47 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{2 \text{ mol Hg}}{1 \text{ mol O}_2} \times \frac{200.59 \text{ g Hg}}{1 \text{ mol Hg}} = 81.11
$$

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

3.17 First, determine the limiting reactant by calculating the moles of $\text{AlCl}_3$ that would be obtained if $\text{Al}$ and $\text{HCl}$ were totally consumed:

$$
0.15 \text{ mol Al} \times \frac{2 \text{ mol AlCl}_3}{2 \text{ mol Al}} = 0.150 \text{ mol AlCl}_3
$$

$$
0.35 \text{ mol HCl} \times \frac{2 \text{ mol AlCl}_3}{6 \text{ mol HCl}} = 0.1166 \text{ mol AlCl}_3
$$

Because the HCl produces the smaller amount of $\text{AlCl}_3$, the reaction will stop when HCl is totally consumed but before the Al is consumed. The limiting reactant is therefore HCl. The amount of $\text{AlCl}_3$ produced must be 0.1166, or 0.12 mol.

3.18 First, determine the limiting reactant by calculating the moles of $\text{ZnS}$ produced by totally consuming Zn and $\text{S}_8$:

$$
7.36 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.39 \text{ g Zn}} \times \frac{8 \text{ mol ZnS}}{8 \text{ mol Zn}} = 0.11256 \text{ mol ZnS}
$$

$$
6.45 \text{ g S}_8 \times \frac{1 \text{ mol S}_8}{256.56 \text{ g S}_8} \times \frac{8 \text{ mol ZnS}}{1 \text{ mol S}_8} = 0.2011 \text{ mol ZnS}
$$

The reaction will stop when Zn is totally consumed; $\text{S}_8$ is in excess, and not all of it is converted to $\text{ZnS}$. The limiting reactant is therefore Zn. Now convert the moles of $\text{ZnS}$ obtained from the Zn to grams of $\text{ZnS}$:

$$
0.11256 \text{ mol ZnS} \times \frac{97.46 \text{ g ZnS}}{1 \text{ mol ZnS}} = 10.97 = 11.0 \text{ g ZnS}
$$
3.19 First, write the balanced equation:

\[ \text{CH}_3\text{OH} + \text{CO} \rightarrow \text{HC}_2\text{H}_3\text{O}_2 \]

Convert grams of each reactant to moles of acetic acid:

\[ 15.0 \text{ g CH}_3\text{OH} \times \frac{1 \text{ mol CH}_3\text{OH}}{32.04 \text{ g CH}_3\text{OH}} \times \frac{1 \text{ mol HC}_2\text{H}_3\text{O}_2}{1 \text{ mol CH}_3\text{OH}} = 0.4681 \text{ mol HC}_2\text{H}_3\text{O}_2 \]

\[ 10.0 \text{ g CO} \times \frac{1 \text{ mol CO}}{28.01 \text{ g CO}} \times \frac{1 \text{ mol HC}_2\text{H}_3\text{O}_2}{1 \text{ mol CH}_3\text{OH}} = 0.3570 \text{ mol HC}_2\text{H}_3\text{O}_2 \]

Thus, CO is the limiting reactant, and 0.03570 mol HC\textsubscript{2}H\textsubscript{3}O\textsubscript{2} is obtained. The mass of product is

\[ 0.3570 \text{ mol HC}_2\text{H}_3\text{O}_2 \times \frac{60.05 \text{ g HC}_2\text{H}_3\text{O}_2}{1 \text{ mol HC}_2\text{H}_3\text{O}_2} = 21.44 \text{ g HC}_2\text{H}_3\text{O}_2 \]

The percentage yield is:

\[ \frac{19.1 \text{ g actual yield}}{21.44 \text{ g theoretical yield}} \times 100\% = 89.08 \% = 89.1\% \]

\[ \text{Mass percentage} = \frac{0.25359 \text{ g HC}_2\text{H}_3\text{O}_2}{5.00 \text{ g vinegar}} \times 100\% = 5.071 = 5.07\% \]

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)\textsubscript{2} has two OH\textsuperscript{-} ions, so there are 1.5 mol x 2 = 3.0 mol OH\textsuperscript{-} 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\textsubscript{2} and H\textsubscript{2}O, respectively. In this experiment, where all of the carbon and hydrogen do not show up, when you analyze the CO\textsubscript{2} for carbon and H\textsubscript{2}O 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. \( \text{C}_2\text{H}_8\text{O}_2 \) is not an empirical formula because each of the subscripts can be divided by two to obtain a possible empirical formula of \( \text{CH}_4\text{O} \). (The empirical formula is not the smallest integer ratio of subscripts.)

b. \( \text{C}_{1.5}\text{H}_4 \) 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 \( \text{C}_3\text{H}_8 \). (Since the subscript of carbon is the decimal number 1.5, the empirical formula is not the smallest integer 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, 12 g of C, 32 g of \( \text{O}_2 \), and 44 g of \( \text{CO}_2 \) 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 \( \text{O}_2 \) 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 \( \text{CO}_2 \) would form.

g. Incorrect. In this representation, oxygen is being shown as individual atoms of O, not as molecules of \( \text{O}_2 \), 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. \( \text{X}_2\text{(g)} + 2\text{Y(g)} \rightarrow 2\text{XY(g)} \)  
(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 ($6.02 \times 10^{23}$) 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$_2$O are obtained as products. Next, you relate the masses of CO$_2$ and H$_2$O 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 \((\text{HO})_2\), or \(\text{H}_2\text{O}_2\).

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 \(\text{CH}_4\), \(\text{O}_2\), \(\text{CO}_2\), and \(\text{H}_2\text{O}\), 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:

\[
\begin{align*}
\text{CH}_4 & + 2\text{O}_2 & \rightarrow & \text{CO}_2 & + & 2\text{H}_2\text{O} \\
1 \text{ molecule} & + 2 \text{ molecules} & \rightarrow & 1 \text{ molecule} & + & 2 \text{ molecules} \\
1 \text{ mole} & + 2 \text{ moles} & \rightarrow & 1 \text{ mole} & + & 2 \text{ moles} \\
16.0 \text{ g} & + 2 \times 32.0 \text{ g} & \rightarrow & 44.0 \text{ g} & + & 2 \times 18.0 \text{ g}
\end{align*}
\]

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.
CHAPTER 3

Answers to Conceptual Problems

3.13 a. \(3\text{H}_2(g) + \text{N}_2(g) \rightarrow 2\text{NH}_3(g)\)

   b. Since there is no \(\text{H}_2\) 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 \(\text{H}_2\) are required for every molecule of \(\text{N}_2\). Since there are two moles of unreacted \(\text{N}_2\), you would need six additional moles of \(\text{H}_2\) 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 \(\text{O}_2\) available for complete combustion, a reasonable assumption is that some of the products will have fewer oxygen atoms than \(\text{CO}_2\). Therefore, a mixture of products would be obtained in this case, including carbon monoxide (\(\text{CO}\)) and soot (carbon particles).

3.15 a. This answer is unreasonable because \(1.0 \times 10^{-3}\) 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 \(1.80 \times 10^{-10}\) g is too large for one water molecule. (The mass of one water molecule is \(2.99 \times 10^{-23}\) g)

   c. This answer is reasonable because \(3.01 \times 10^{23}\) 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 \(\text{H}_2\) to one mole of \(\text{O}_2\). According to the balanced chemical reaction, every one mole of \(\text{O}_2\) can react with two moles of \(\text{H}_2\). In this case, when 0.5 mole of \(\text{O}_2\) has reacted, all of the \(\text{H}_2\) (one mole) will be consumed, leaving behind 0.5 mole of unreacted \(\text{O}_2\).

   (continued)
c. In this case, you have a ratio of two moles of H\textsubscript{2} to one mole of O\textsubscript{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\textsubscript{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\textsubscript{2} + O\textsubscript{2} → 2H\textsubscript{2}O 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\textsubscript{2}O molecules and two O\textsubscript{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

\[
\frac{27.0 \text{ g Ca}}{40.08 \text{ g Ca}} \times \frac{1 \text{ mol Ca}}{1 \text{ mol Ca}} = 0.674 \text{ mol Ca}
\]

b. The problem here is an incorrect mole ratio. There are two mol K\textsuperscript{+} ions per one mol K\textsubscript{2}SO\textsubscript{4}. The correct calculation is

\[
\frac{2.5 \text{ mol K}_2\text{SO}_4}{1 \text{ mol K}_2\text{SO}_4} \times \frac{2 \text{ mol K}^+ \text{ ions}}{1 \text{ mol K}_2\text{SO}_4} \times \frac{6.022 \times 10^{23} \text{ K}^+ \text{ ions}}{1 \text{ mol K}^+ \text{ ions}} = 3.01 \times 10^{24} \text{ K}^+ \text{ ions}
\]

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

\[
\frac{0.50 \text{ mol Na}}{2 \text{ mol Na}} \times \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol Na}} = 0.50 \text{ mol H}_2\text{O}
\]
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 $5 \text{ mol O}_2 / 2 \text{ mol C}_2\text{H}_2$. This means that, for every two moles of C$_2$H$_2$, 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$_3$OH $= \text{AW of C} + 4(\text{AW of H}) + \text{AW of O}$. 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

$$\text{FW} = 12.01 \text{ amu} + (4 \times 1.008 \text{ amu}) + 16.00 \text{ amu} = 32.042 = 32.0 \text{ amu}$$

b. FW of NO$_3$ $= \text{AW of N} + 3(\text{AW of O})$

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

c. FW of K$_2$CO$_3$ $= 2(\text{AW of K}) + \text{AW of C} + 3(\text{AW of O})$

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

$$= 138.210 = 138 \text{ amu}$$

d. FW of Ni$_3$(PO$_4$)$_2$ $= 3(\text{AW of Ni}) + 2(\text{AW of P}) + 8(\text{AW of O})$

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

$$= 366.040 = 366 \text{ amu}$$

3.22 a. FW of HC$_2$H$_3$O$_2$ $= 2(\text{AW of C}) + 4(\text{AW of H}) + 2(\text{AW of O})$

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

$$= 60.052 = 60.1 \text{ amu}$$

(continued)
CALCULATIONS WITH CHEMICAL FORMULAS AND EQUATIONS

b. FW of PCl

\[ \text{FW of PCl}_5 = \text{AW of P} + 5(\text{AW of Cl}) \]
\[ = 30.97 \text{ amu} + (5 \times 35.45 \text{ amu}) \]
\[ = 208.220 = 208 \text{ amu} \]

c. FW of K₂SO₄

\[ \text{FW of K}_2\text{SO}_4 = 2(\text{AW of K}) + \text{AW of S} + 4(\text{AW of O}) \]
\[ = (2 \times 39.10 \text{ amu}) + 32.07 \text{ amu} + (4 \times 16.00 \text{ amu}) \]
\[ = 174.270 = 174 \text{ amu (3 s.f.)} \]

d. FW of Ca(OH)₂

\[ \text{FW of Ca(OH)}_2 = \text{AW of Ca} + 2(\text{AW of H}) + 2(\text{AW of O}) \]
\[ = 40.08 \text{ amu} + (2 \times 1.008 \text{ amu}) + (2 \times 16.00 \text{ amu}) \]
\[ = 74.096 = 74.1 \text{ amu (3 s.f.)} \]

3.23 a  SO₂

\[ 1 \times \text{AW of S} = 32.07 \text{ amu} \]
\[ 2 \times \text{AW of O} = 2 \times 16.00 = 32.00 \text{ amu} \]
\[ \text{MW of NO}_2 = 64.07 \text{ amu} = 64.1 \text{ amu (3 s.f.)} \]

b. PCl₃

\[ 1 \times \text{AW of P} = 30.97 \text{ amu} \]
\[ 3 \times \text{AW of Cl} = 3 \times 35.45 = 106.35 \text{ amu} \]
\[ \text{MW of PCl}_3 = 137.32 \text{ amu} = 137 \text{ amu (3 s.f.)} \]

3.24 a  HNO₂

\[ 1 \times \text{AW of H} = 1.008 \text{ amu} \]
\[ 1 \times \text{AW of N} = 14.01 \text{ amu} \]
\[ 2 \times \text{AW of O} = 2 \times 16.00 = 32.00 \text{ amu} \]
\[ \text{MW of HNO}_2 = 47.018 \text{ amu} = 47.0 \text{ amu (3 s.f.)} \]

b. CO

\[ 1 \times \text{AW of C} = 12.01 \text{ amu} \]
\[ 1 \times \text{AW of O} = 16.00 \text{ amu} \]
\[ \text{MW of CO} = 28.01 \text{ amu} = 28.0 \text{ amu (3 s.f.)} \]

3.25 First, find the formula weight of NH₄NO₃ by adding the respective atomic weights. Then convert it to the molar mass:

\[ \text{FW of NH}_4\text{NO}_3 = 2(\text{AW of N}) + 4(\text{AW of H}) + 3(\text{AW of O}) \]
\[ = (2 \times 14.01 \text{ amu}) + (4 \times 1.008 \text{ amu}) + (3 \times 16.00 \text{ amu}) \]
\[ = 80.052 \text{ amu} \]

The molar mass of NH₄NO₃ = 80.05 g/mol.
3.26 First, find the formula weight of $\text{H}_3\text{PO}_4$ by adding the respective atomic weights. Then convert it to the molar mass:

$$\text{FW of } \text{H}_3\text{PO}_4 = 3(\text{AW of H}) + \text{AW of P} + 4(\text{AW of O})$$

$$= (3 \times 1.008) + 30.97 + (4 \times 16.00)$$

$$= 97.994 \text{ amu}$$

The molar mass of $\text{H}_3\text{PO}_4 = 97.99 \text{ g/mol}$. 

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 $6.022 \times 10^{23}$ Na atoms, we calculate

$$\text{Mass of one Na atom} = \frac{22.99 \text{ g/mol}}{6.022 \times 10^{23} \text{ atom/mol}} = 3.8177 \times 10^{-23} \text{ g/atom}$$

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 $6.022 \times 10^{23}$ N atoms, we calculate

$$\text{Mass of one N atom} = \frac{14.01 \text{ g/mol}}{6.022 \times 10^{23} \text{ atom/mol}} = 2.3264 \times 10^{-23} \text{ g/atom}$$

c. The formula weight of CH$_3$Cl = [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$_3$Cl molecules equals $6.022 \times 10^{23}$ CH$_3$Cl molecules, we calculate

$$\text{Mass of one CH}_3\text{Cl molecule} = \frac{50.48 \text{ g/mol}}{6.022 \times 10^{23} \text{ molecules/mol}} = 8.3826 \times 10^{-23} \text{ g/molecule}$$

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 $6.022 \times 10^{23}$ Hg(NO$_3$)$_2$ formula units, we calculate

$$\text{Mass of one Hg(NO}_3)_2 = \frac{324.61 \text{ g/mol}}{6.022 \times 10^{23} \text{ units/mol}} = 5.3904 \times 10^{-22} \text{ g/unit}$$

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. Mass of one Fe atom = \(\frac{55.85 \text{ g/mol}}{6.022 \times 10^{23} \text{ atom/mol}} = 9.2743 \times 10^{-23} \text{ g/atom}\) (continued)
b. Mass of one F atom = \( \frac{19.00 \text{ g/mol}}{6.022 \times 10^{23} \text{ atom/mol}} \) = \( 3.1551 \times 10^{-23} \text{ g/atom} \)

c. Mass of one N\(_2\)O molecule = \( \frac{44.02 \text{ g/mol}}{6.022 \times 10^{23} \text{ molecules/mol}} \) = \( 7.3099 \times 10^{-23} \text{ g/molecule} \)

d. Mass of one Al(OH)\(_3\) unit = \( \frac{78.00 \text{ g/mol}}{6.022 \times 10^{23} \text{ unit/mol}} \) = \( 1.2953 \times 10^{-22} \text{ g/unit} \)

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

FW of \((\text{CH}_3\text{CH}_2)_2\text{O}\) = \((4 \times 12.01 \text{ amu}) + (10 \times 1.008 \text{ amu}) + 16.00 \text{ amu}\)

= 74.12 \text{ amu}

Mass of one \((\text{CH}_3\text{CH}_2)_2\text{O}\) molecule = \( \frac{74.12 \text{ g/mol}}{6.022 \times 10^{23} \text{ molecules/mol}} \) = \( 1.2308 \times 10^{-22} \text{ g/molecule} \)

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

FW of glycerol = \((3 \times 12.01 \text{ amu}) + (8 \times 1.008 \text{ amu}) + (3 \times 16.00 \text{ amu})\)

= 92.09 \text{ amu}

Mass of one glycerol molecule = \( \frac{92.09 \text{ g/mol}}{6.022 \times 10^{23} \text{ molecules/mol}} \) = \( 1.529 \times 10^{-22} \text{ g/molecule} \)

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.

\[ a. \quad 0.15 \text{ mol Na} \times \frac{22.99 \text{ g}}{1\text{ mol Na}} = 3.448 = 3.4 \text{ g Na} \]

\[ b. \quad 0.594 \text{ mol S} \times \frac{32.07 \text{ g S}}{1\text{ mol S}} = 19.04 = 19.0 \text{ g S} \]

(continued)
c. Using molar mass $= 84.93$ g/mol for $\text{CH}_2\text{Cl}_2$, we obtain

$$2.78 \text{ mol CH}_2\text{Cl}_2 \times \frac{84.93 \text{ g CH}_2\text{Cl}_2}{1 \text{ mol CH}_2\text{Cl}_2} = 236.1 = 236 \text{ g CH}_2\text{Cl}_2$$

d. Using molar mass $= 68.14$ g/mol for $(\text{NH}_4)_2\text{S}$, we obtain

$$38 \text{ mol (NH}_4)_2\text{S} \times \frac{68.14 \text{ g (NH}_4)_2\text{S}}{1 \text{ mol (NH}_4)_2\text{S}} = 2.58 \times 10^3 = 2.6 \times 10^3 \text{ g (NH}_4)_2\text{S}$$

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

a. $0.205 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 11.44 = 11.4 \text{ g Fe}$

b. $0.79 \text{ mol F} \times \frac{19.00 \text{ g F}}{1 \text{ mol F}} = 15.01 = 15 \text{ g F}$

c. Using molar mass $= 44.01$ g/mol for $\text{CO}_2$, we obtain

$$5.8 \text{ mol CO}_2 \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 255.2 = 2.6 \times 10^2 \text{ g CO}_2$$

d. Using molar mass $194.20$ g/mol for $\text{K}_2\text{CrO}_4$, we obtain

$$48.1 \text{ mol K}_2\text{CrO}_4 \times \frac{194.20 \text{ g K}_2\text{CrO}_4}{1 \text{ mol K}_2\text{CrO}_4} = 9341.02 = 9.34 \times 10^3 \text{ g K}_2\text{CrO}_4$$

3.33 First, find the molar mass of $\text{H}_3\text{BO}_3$: $(3 \times 1.008 \text{ amu}) + 10.81 \text{ amu} + (3 \times 16.00 \text{ amu}) = 61.83$. Therefore, the molar mass of $\text{H}_3\text{BO}_3 = 61.83$ g/mol. The mass of $\text{H}_3\text{BO}_3$ is calculated as follows:

$$0.543 \text{ mol H}_3\text{BO}_3 \times \frac{61.83 \text{ g H}_3\text{BO}_3}{1 \text{ mol H}_3\text{BO}_3} = 33.57 = 33.6 \text{ g H}_3\text{BO}_3$$

3.34 First, find the molar mass of $\text{CS}_2$: $12.01 \text{ amu} + (2 \times 32.07 \text{ amu}) = 76.15$ amu. Therefore, the molar mass of $\text{CS}_2 = 76.15$ g/mol. The mass of $\text{CS}_2$ is calculated as follows:

$$0.0205 \text{ mol CS}_2 \times \frac{76.15 \text{ g CS}_2}{1 \text{ mol CS}_2} = 1.5610 = 1.56 \text{ g CS}_2$$
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. \[ \text{2.86 g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.2381 \text{ mol C} \]

b. \[ \text{7.05 g Cl}_2 \times \frac{1 \text{ mol Cl}_2}{70.90 \text{ g Cl}_2} = 0.09943 \text{ mol Cl}_2 \]

c. The molar mass of \( \text{C}_4\text{H}_{10} \) = \( (4 \times 12.01) + (10 \times 1.008) \) = 58.12 g/mol \( \text{C}_4\text{H}_{10}/\text{mol C}_4\text{H}_{10} \). The mass of \( \text{C}_4\text{H}_{10} \) is calculated as follows:

\[ \text{76 g C}_4\text{H}_{10} \times \frac{1 \text{ mol C}_4\text{H}_{10}}{58.12 \text{ g C}_4\text{H}_{10}} = 1.307 \text{ mol C}_4\text{H}_{10} \]

d. The molar mass of \( \text{Al}_2(\text{CO}_3)_3 \) = \( (2 \times 26.98) + (3 \times 12.01) + (9 \times 16.00) \) = 233.99 g/mol \( \text{Al}_2(\text{CO}_3)_3 \). The mass of \( \text{Al}_2(\text{CO}_3)_3 \) is calculated as follows:

\[ \text{26.2 g Al}_2(\text{CO}_3)_3 \times \frac{1 \text{ mol Al}_2(\text{CO}_3)_3}{233.99 \text{ g Al}_2(\text{CO}_3)_3} = 0.1119 \text{ mol Al}_2(\text{CO}_3)_3 \]

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. \[ \text{2.57 g As} \times \frac{1 \text{ mol As}}{74.92 \text{ g As}} = 0.03430 \text{ mol As} \]

b. \[ \text{7.83 g S}_8 \times \frac{1 \text{ mol S}_8}{256.56 \text{ g S}_8} = 0.03051 \text{ mol S}_8 \]

c. The molar mass of \( \text{N}_2\text{H}_4 \) = \( (2 \times 14.01) + (4 \times 1.008) \) = 32.052 g/mol \( \text{N}_2\text{H}_4 \). The mass of \( \text{N}_2\text{H}_4 \) is calculated as follows:

\[ \text{36.5 g} \times \frac{1 \text{ mol N}_2\text{H}_4}{32.052 \text{ g N}_2\text{H}_4} = 1.1387 \text{ g} \]

d. The molar mass of \( \text{Al}_2(\text{SO}_4)_3 \) = \( (2 \times 26.98) + (3 \times 32.07) + (12 \times 16.00) \) = 342.17 g/mol \( \text{Al}_2(\text{SO}_4)_3 \). The mass of \( \text{Al}_2(\text{SO}_4)_3 \) is calculated as follows:

\[ \text{227 g Al}_2(\text{SO}_4)_3 \times \frac{1 \text{ mol Al}_2(\text{SO}_4)_3}{342.17 \text{ g Al}_2(\text{SO}_4)_3} = 0.6634 \text{ mol Al}_2(\text{SO}_4)_3 \]
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 CaSO₄ is 136.15 g/mol. Use this to convert the mass of CaSO₄ to moles:

\[
0.791 \text{ g CaSO}_4 \times \frac{1 \text{ mol CaSO}_4}{136.15 \text{ g CaSO}_4} = 5.811 \times 10^{-3} = 5.81 \times 10^{-3} \text{ mol CaSO}_4
\]

Calculate the molecular weight of water: (2 x 1.008 amu) + 16.00 amu = 18.02 amu. Therefore, the molar mass of H₂O equals 18.02 g/mol. Use this to convert the rest of the sample to moles of water:

\[
0.209 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 1.159 \times 10^{-2} = 1.16 \times 10^{-2} \text{ mol H}_2\text{O}
\]

Because 0.01159 mol is about twice 0.005811 mol, both numbers of moles are consistent with the formula, CaSO₄•2H₂O.

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 CuSO₄ is 159.62 g/mol. From the previous problem, the molar mass of H₂O is 18.02 g/mol. Use this to convert the rest of the sample to moles of water:

\[
0.558 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 0.03096 \text{ mol H}_2\text{O}
\]

Calculate the moles of CuSO₄ to be able to compare relative molar amounts of CuSO₄ and H₂O. Then, divide the moles of H₂O by the moles of CuSO₄:

\[
0.989 \text{ g CuSO}_4 \times \frac{1 \text{ mol CuSO}_4}{159.62 \text{ g CuSO}_4} = 0.006196 \text{ mol CuSO}_4
\]

\[
0.03096 \text{ mol H}_2\text{O} \div 0.006196 \text{ mol CuSO}_4 = 4.99/1, \text{ or about 5:1 (consistent with CuSO}_4\cdot5\text{H}_2\text{O)}
\]

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 x 10²³ atoms, so

a. No. Li atoms = 8.21 g Li x \(\frac{6.022 \times 10^{23} \text{ atoms}}{6.941 \text{ g Li}}\) = 7.122 x 10²³ atoms

b. No. Br atoms = 32.0 g Br₂ x \(\frac{2 \times 6.022 \times 10^{23} \text{ atoms}}{(2 \times 79.90) \text{ g Br}_2}\) = 2.412 x 10²³ atoms

(continued)
c. No. NH₃ molecules = 45 g NH₃ x $\frac{6.022 \times 10^{23} \text{molecules}}{17.03 \text{g NH}_3} = 1.59 \times 10^{24} \text{molecules}$

d. No. PbCrO₄ units = 201 g PbCrO₄ x $\frac{6.022 \times 10^{23} \text{units}}{323.2 \text{g PbCrO}_4} = 3.745 \times 10^{23} \text{units}$

e. No. SO₄²⁻ ions = 14.3 g Cr₂(SO₄)₃ x $\frac{3 \times 6.022 \times 10^{23} \text{ions}}{392.21 \text{g Cr}_2(\text{SO}_4)_3} = 6.587 \times 10^{22} \text{ions}$

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 $6.022 \times 10^{23} \text{atoms}$, so

a. No. Al atoms = 25.7 g Al x $\frac{6.022 \times 10^{23} \text{atoms}}{26.98 \text{g Al}} = 5.736 \times 10^{23} \text{atoms}$

b. No. I atoms = 8.71 g I₂ x $\frac{2 \times 6.022 \times 10^{23} \text{atoms}}{2 \times 126.90 \text{g I}_2} = 4.133 \times 10^{22} \text{atoms}$

c. No. N₂O₅ molecules = 14.9 g N₂O₅ x $\frac{6.022 \times 10^{23} \text{molecules}}{108.02 \text{g N}_2\text{O}_5} = 8.306 \times 10^{22} \text{molecules}$

d. No. NaClO₄ units = 3.31 g NaClO₄ x $\frac{6.022 \times 10^{23} \text{units}}{122.44 \text{g NaClO}_4} = 1.628 \times 10^{22} \text{units}$

e. No. Ca²⁺ ions = 4.71 g Ca₃(PO₄)₂ x $\frac{3 \times 6.022 \times 10^{23} \text{ions}}{310.18 \text{g Ca}_3(\text{PO}_4)_2} = 2.743 \times 10^{22} \text{ions}$

3.41 Calculate the molecular weight of CCl₄: 12.01 amu + (4 x 35.45 amu) = 153.81 amu. Use this and Avogadro’s number to express it as 153.81 g/Nₐ to calculate the number of molecules:

$7.58 \text{mg CCl}_4 \times \frac{1 \text{g}}{1000 \text{mg}} \times \frac{6.022 \times 10^{23} \text{molecules}}{153.81 \text{g CCl}_4} = 2.968 \times 10^{19}$

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

\[
5.88 \text{ mg } \text{ClF}_3 \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{92.45 \text{ g } \text{ClF}_3} = 3.83 \times 10^{19}
\]

\[
= 3.83 \times 10^{19} \text{ molecules}
\]

3.43 Mass percentage carbon = \(\frac{\text{mass of C in sample}}{\text{mass of sample}}\) x 100%

\[
\text{Percent carbon} = \frac{1.584 \text{ g}}{1.836 \text{ g}} \times 100\% = 86.27\% = 86.27\%
\]

3.44 Mass percentage alcohol = \(\frac{\text{mass of alcohol in solution}}{\text{mass of sample}}\) x 100%

\[
\text{Percent alcohol} = \frac{3.98 \text{ g}}{6.01 \text{ g}} \times 100\% = 66.22\% = 66.2\%
\]

3.45 Mass percentage phosphorus oxychloride = \(\frac{\text{mass of POCl}_3 \text{ in sample}}{\text{mass of sample}}\) x 100%

\[
\text{Percent POCl}_3 = \frac{1.72 \text{ mg}}{8.53 \text{ mg}} \times 100\% = 20.16\% = 20.2\%
\]

3.46 Mass percentage sulfur = \(\frac{\text{mass of S in sample}}{\text{mass of sample}}\) x 100%

\[
\text{Percent sulfur} = \frac{1.64 \text{ mg}}{3.17 \text{ mg}} \times 100\% = 51.73\% = 51.7\%
\]
3.47 Start with the definition for percentage nitrogen, and rearrange this equation to find the mass of N in the fertilizer.

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

\[
\text{Mass N} = \frac{\text{mass \% N}}{100\%} \times \text{mass of fertilizer} = \frac{14.0\%}{100\%} \times 4.15 \text{ kg} = 0.5810
\]

\[
= 0.581 \text{ kg N}
\]

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.

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

\[
\text{Mass Br} = \frac{\text{mass \% Br}}{100\%} \times \text{mass seawater} = \frac{0.0065\%}{100\%} \times 1025 \text{ g} = 0.0666
\]

\[
= 0.067 \text{ g Br}
\]

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

\[
0.0898 \text{ mol Al} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} = 2.422 \text{ g Al}
\]

\[
0.0381 \text{ mol Mg} \times \frac{24.31 \text{ g Mg}}{1 \text{ mol Mg}} = 0.9262 \text{ g Mg}
\]

Percent Al = \[
\frac{\text{mass of Al}}{\text{mass of alloy}} \times 100\% = \frac{2.422 \text{ g Al}}{3.349 \text{ g alloy}} \times 100\% = 72.34 = 72.3\% \text{ Al}
\]

Percent Mg = \[
\frac{\text{mass of Mg}}{\text{mass of alloy}} \times 100\% = \frac{0.9262 \text{ g Mg}}{3.349 \text{ g alloy}} \times 100\% = 27.655 = 27.7\% \text{ Mg}
\]
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.

\[
\begin{align*}
0.0856 \text{ mol Ne} & \times \frac{20.18 \text{ g Ne}}{1 \text{ mol Ne}} = 1.727 \text{ g Ne} \\
0.0254 \text{ mol Kr} & \times \frac{83.80 \text{ g Kr}}{1 \text{ mol Kr}} = 2.129 \text{ g Kr}
\end{align*}
\]

\[
\begin{align*}
\text{Percent Ne} = \frac{\text{mass of Ne}}{\text{mass of mix}} \times 100\% = \frac{1.727 \text{ g Ne}}{3.856 \text{ g mix}} \times 100\% = 44.79 \% = 44.8\% \text{ Ne}
\end{align*}
\]

\[
\begin{align*}
\text{Percent Kr} = \frac{\text{mass of Kr}}{\text{mass of mix}} \times 100\% = \frac{2.129 \text{ g Kr}}{3.856 \text{ g mix}} \times 100\% = 55.21 \% = 55.2\% \text{ Kr}
\end{align*}
\]

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

\[
\text{Percent O} = 100.000\% - 42.878\% \text{C} = 57.122 = 57.1\%
\]

b. \[
\text{Percent C} = \frac{\text{mass of C}}{\text{mass of CO}_2} \times 100\% = \frac{12.01 \text{ g C}}{44.01 \text{ g CO}_2} \times 100\% = 27.289 = 27.3\%
\]

\[
\text{Percent O} = 100.000\% - 27.289\% \text{ C} = 72.711 = 72.7\%
\]

c. \[
\text{Percent Na} = \frac{\text{mass of Na}}{\text{mass of NaH}_2\text{PO}_4} \times 100\% = \frac{22.99 \text{ g Na}}{119.98 \text{ g NaH}_2\text{PO}_4} \times 100\% = 19.161 = 19.2\%
\]

\[
\text{Percent H} = \frac{\text{mass of H}}{\text{mass of NaH}_2\text{PO}_4} \times 100\% = \frac{2.016 \text{ g H}}{119.98 \text{ g NaH}_2\text{PO}_4} \times 100\% = 1.6802 = 1.68\%
\]

\[
\text{Percent P} = \frac{\text{mass of P}}{\text{mass of NaH}_2\text{PO}_4} \times 100\% = \frac{30.97 \text{ g P}}{119.98 \text{ g NaH}_2\text{PO}_4} \times 100\% = 25.812 = 25.8\%
\]

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

(continued)
d. Percent Co = \( \frac{\text{mass of Co}}{\text{mass of Co(NO}_3\text{)_2}} \) = \( \frac{58.93 \text{ g Co}}{182.95 \text{ g Co(NO}_3\text{)_2}} \) x 100%  
\[ = 32.211 = 32.2\% \]

Percent N = \( \frac{\text{mass of N}}{\text{mass of Co(NO}_3\text{)_2}} \) = \( \frac{2 \times 14.01 \text{ g N}}{182.95 \text{ g Co(NO}_3\text{)_2}} \) x 100%  
\[ = 15.316 = 15.3\% \]

Percent O = 100.000% - (32.211 + 15.316) = 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. Percent N = \( \frac{\text{mass of N}}{\text{mass of NO}} \) = \( \frac{14.01 \text{ g N}}{30.01 \text{ g NO}} \) x 100% = 46.684 = 46.7%

Percent O = 100.000% - 46.684% N = 53.316 = 53.3%

b. Percent H = \( \frac{\text{mass of H}}{\text{mass of H}_2\text{O}_2} \) = \( \frac{2.016 \text{ g H}}{34.02 \text{ g H}_2\text{O}_2} \) x 100% = 5.9259 = 5.93%

Percent O = 100.000% - 5.2959% H = 94.07 = 94.1%

c. Percent K = \( \frac{\text{mass of K}}{\text{mass of KClO}_4} \) = \( \frac{39.10 \text{ g K}}{138.55 \text{ g KClO}_4} \) x 100% = 28.221 = 28.2%

Percent Cl = \( \frac{\text{mass of Cl}}{\text{mass of KClO}_4} \) = \( \frac{35.45 \text{ g Cl}}{138.55 \text{ g KClO}_4} \) x 100% = 25.586 = 25.6%

Percent O = 100.000% - (28.221 + 25.586) = 46.193 = 46.2%

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

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

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:

$$\text{Mol C (glucose)} = 6.01 \text{ g} \times \frac{1 \text{ mol}}{180.2 \text{ g}} \times \frac{6 \text{ mol C}}{1 \text{ mol glucose}} = 0.200 \text{ mol}$$

$$\text{Mol C (ethanol)} = 5.85 \text{ g} \times \frac{1 \text{ mol}}{46.07 \text{ g}} \times \frac{2 \text{ mol C}}{1 \text{ mol ethanol}} = 0.254 \text{ mol (more C)}$$

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.

$$\text{Mol S (CaSO}_4\text{)} = 40.8 \text{ g} \times \frac{1 \text{ mol CaSO}_4}{136.15 \text{ g}} \times \frac{1 \text{ mol S}}{1 \text{ mol CaSO}_4} = 0.2997 \text{ mol (more S)}$$

$$\text{Mol S (Na}_2\text{SO}_3\text{)} = 35.2 \text{ g} \times \frac{1 \text{ mol Na}_2\text{SO}_3}{126.05 \text{ g}} \times \frac{1 \text{ mol S}}{1 \text{ mol Na}_2\text{SO}_3} = 0.2793 \text{ mol}$$
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$_2$O by the molar mass of 2H and the reciprocal of the molar mass of H$_2$O. Then, use the masses to calculate the mass percentages. Calculate O by difference.

\[
9.06 \text{ mg } \text{CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 2.47 \text{ mg C}
\]

\[
5.58 \text{ mg } \text{H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ H}}{1 \text{ H}_2\text{O}} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.6243 \text{ mg H}
\]

Mass O = 6.38 mg - (2.472 + 0.6243) = 3.284 mg O

Percent C = (2.472 mg C/6.38 mg glycol) x 100% = 38.74 = 38.7%

Percent H = (0.6243 mg H/6.38 mg glycol) x 100% = 9.785 = 9.79%

Percent O = (3.284 mg O/6.38 mg glycol) x 100% = 51.47 = 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$_2$O by the molar mass of 2H and the reciprocal of the molar mass of H$_2$O. Then, use the masses to calculate the mass percentages. Calculate O by difference.

\[
14.67 \text{ mg } \text{CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 4.0033 \text{ mg C}
\]

\[
3.01 \text{ mg } \text{H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ H}}{1 \text{ H}_2\text{O}} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.3368 \text{ mg H}
\]

Mass O = 5.23 mg - (4.0033 + 0.3368) = 0.8899 mg O

Percent C = (4.0033 mg/5.23 mg) x 100% = 76.54 = 76.5%

Percent H = (0.3368 mg/5.23 mg) x 100% = 6.439 = 6.44%

Percent O = (0.8899 mg/5.23 mg) x 100% = 17.0 = 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.

\[
\text{Mol Os} = \frac{2.16 \text{ g Os}}{190.2 \text{ g Os}} = 0.01136 \text{ mol (smaller number)}
\]

\[
\text{Mol O} = \frac{(2.89 - 2.16) \text{ g O}}{16.00 \text{ g O}} = 0.0456 \text{ mol}
\]

Integer for Os = \(0.01136 \div 0.01136 = 1.000\)

Integer for O = \(0.0456 \div 0.01136 = 4.01\)

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.

\[
\text{Mol W} = \frac{4.23 \text{ g W}}{183.85 \text{ g W}} = 0.02301 \text{ mol (smaller number)}
\]

\[
\text{Mol O} = \frac{(5.34 - 4.23) \text{ g O}}{16.00 \text{ g O}} = 0.06938 \text{ mol}
\]

Integer for W = \(0.02301 \div 0.02301 = 1.000\)

Integer for O = \(0.06938 \div 0.02301 = 3.015\)

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.

\[
\text{Mol K} = \frac{39.6 \text{ g K}}{39.10 \text{ g K}} = 1.013 \text{ mol}
\]

\[
\text{Mol Mn} = \frac{29.7 \text{ g Mn}}{54.94 \text{ g Mn}} = 0.5078 \text{ mol (smallest number)}
\]

\[
\text{Mol O} = \frac{32.5 \text{ g O}}{16.00 \text{ g O}} = 2.031 \text{ mol}
\]

(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 = $1.013 \div 0.5078 = 1.998$, or 2

Integer for Mn = $0.5078 \div 0.5078 = 1.000$, or 1

Integer for O = $2.031 \div 0.5078 = 3.999$, 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.

Mol C = $65.4 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 5.445 \text{ mol}$

Mol H = $5.5 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 5.46 \text{ mol}$

Mol O = $29.1 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.819 \text{ mol} \text{ (smallest number)}$

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

Integer for C = $5.445 \div 1.819 = 2.99$, or 3

Integer for H = $5.46 \div 1.819 = 3.0$, or 3

Integer for O = $1.819 \div 1.819 = 1.00$, 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.

Mol C = $50.0 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.163 \text{ mol}$

Mol H = $5.6 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 5.56 \text{ mol}$

(continued)
CHAPTER 3

\[ \text{Mol O} = 44.0 \text{ g O} \times \frac{1\text{mol O}}{16.00 \text{g O}} = 2.775 \text{ mol (smallest number)} \]

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 \div 2.775 = 1.50, or 1.5
- Tentative integer for H = 5.56 \div 2.775 = 2.00, or 2
- Tentative integer for O = 2.775 \div 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 \times 1.5 = 3
- H: 2 \times 2 = 4
- O: 2 \times 1 = 2

The empirical formula is thus \( \text{C}_3\text{H}_4\text{O}_2 \).

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.

\[ \text{Mol C} = 34.6 \text{ g C} \times \frac{1\text{mol C}}{12.01 \text{g C}} = 2.881 \text{ mol (smallest number)} \]

\[ \text{Mol H} = 3.9 \text{ g H} \times \frac{1\text{mol H}}{1.008 \text{g H}} = 3.87 \text{ mol} \]

\[ \text{Mol O} = 61.5 \text{ g O} \times \frac{1\text{mol O}}{16.00 \text{g O}} = 3.844 \text{ mol} \]

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 \div 2.881 = 1.00, or 1
- Tentative integer for H = 3.87 \div 2.881 = 1.34, or 1-1/3
- Tentative integer for O = 3.844 \div 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 C₃H₄O₄.

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:

\[
\begin{align*}
C \text{ mol} &= \frac{92.25 \text{ g C}}{12.01 \text{ g C}} = 7.68109 \text{ mol C} \\
H \text{ mol} &= \frac{7.75 \text{ g H}}{1.008 \text{ g H}} = 7.688 \text{ mol H}
\end{align*}
\]

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:

For 52.03 : \( n = \frac{52.03 \text{ amu}}{13.018 \text{ amu}} = 3.9968 \), or 4

For 78.05 : \( n = \frac{78.05 \text{ amu}}{13.018 \text{ amu}} = 5.9955 \), or 6

The molecular formulas are: for 52.03, (CH)₄ or C₄H₄; and for 78.05, (CH)₆ or C₆H₆.

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:

\[
\begin{align*}
85.62 \text{ g C} &= x \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 7.12905 \text{ mol C} \\
14.38 \text{ g H} &= x \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 14.26 \text{ mol H}
\end{align*}
\]

(continued)
Divide both mole numbers by the smaller one:

For C: \( \frac{7.129 \text{ mol}}{7.129 \text{ mol}} = 1.00 \)

For H: \( \frac{14.26 \text{ mol}}{7.129 \text{ mol}} = 2.0002 \)

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:

For 28.03: \( n = \frac{28.03 \text{ amu}}{14.026 \text{ amu}} = 1.9984 \), or 2

For 56.06: \( n = \frac{56.06 \text{ amu}}{14.026 \text{ amu}} = 3.9968 \), 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.

\[
\text{Formula weight} = (2 \times 12.01 \text{ amu}) + (6 \times 1.008 \text{ amu}) + 14.01 \text{ amu} = 44.08 \text{ amu}
\]

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 \( \{90 + 87 \div 2 \} \), this quotient is

\[ 88.5 \text{ amu} \div 44.1 \text{ amu} = 2.006, \text{ or } 2 \]

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

\[ 28 \text{ amu} \div 13.83 \text{ amu} = 2.02 \]

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.

\[
\begin{align*}
\text{Mol C} &= \frac{26.7 \text{ g C}}{12.01 \text{ g C}} = 2.223 \text{ mol} \\
\text{Mol H} &= \frac{2.2 \text{ g H}}{1.008 \text{ g H}} = 2.18 \text{ mol (smallest number)} \\
\text{Mol O} &= \frac{71.1 \text{ g O}}{16.00 \text{ g O}} = 4.443 \text{ mol}
\end{align*}
\]

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

\[
\begin{align*}
\text{Integer for C} &= 2.223 \div 2.18 = 1.02, \text{ or } 1 \\
\text{Integer for H} &= 2.18 \div 2.18 = 1.00, \text{ or } 1 \\
\text{Integer for O} &= 4.443 \div 2.18 = 2.038, \text{ or } 2
\end{align*}
\]

The empirical formula is thus CHO₂. The formula weight corresponding to this formula may be found by adding the respective atomic weights:

\[
\text{Formula weight} = 12.01 \text{ amu} + 1.008 \text{ amu} + (2 \times 16.00 \text{ amu}) = 45.02 \text{ amu}
\]

Dividing the molecular weight by the formula weight gives the number of times the CHO₂ unit occurs in the molecule. Because the molecular weight is 90 amu, this quotient is

\[
90 \text{ amu} \div 45.02 \text{ amu} = 2.00, \text{ or } 2
\]

The molecular formula is thus (CHO₂)₂, or C₂H₂O₄.

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.

\[
\begin{align*}
\text{Mol C} &= \frac{49.3 \text{ g C}}{12.01 \text{ g C}} = 4.105 \text{ mol} \\
\text{Mol H} &= \frac{6.9 \text{ g H}}{1.008 \text{ g H}} = 6.85 \text{ mol}
\end{align*}
\]

(continued)
\[
\text{Mol O} = 43.8 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.738 \text{ mol (smallest number)}
\]

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 = \(4.105 \div 2.738 = 1.499\), or 1.5

Tentative integer for H = \(6.85 \div 2.738 = 2.50\), or 2.5

Tentative integer for O = \(2.738 \div 2.738 = 1.000\), 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 \(\text{C}_3\text{H}_5\text{O}_2\). The formula weight corresponding to this formula may be found by adding the respective atomic weights:

\[
\text{Formula weight} = (3 \times 12.01 \text{ amu}) + (5 \times 1.008 \text{ amu}) + (2 \times 16.00 \text{ amu})
\]
\[
= 73.1 \text{ amu}
\]

Dividing the molecular weight by the formula weight gives the number of times the \(\text{CHO}_2\) unit occurs in the molecule. Because the molecular weight is 146 amu, this quotient is

\[
146 \text{ amu} \div 73.1 \text{ amu} = 2.00, \text{ or } 2
\]

The molecular formula is thus \((\text{C}_3\text{H}_5\text{O}_2)_2\), or \(\text{C}_6\text{H}_{10}\text{O}_4\).
3.72 \[2H_2S \quad + \quad 3O_2 \quad \rightarrow \quad 2SO_2 \quad + \quad 2H_2O\]

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 \(H_2O\)

\(2 \times 34.06 \text{ g } H_2S \quad + \quad 3 \times 32.00 \text{ g } O_2 \quad \rightarrow \quad 2 \times 64.06 \text{ g } CO_2 \quad + \quad 2 \times 18.016 \text{ g } H_2O\)

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 \(H_2O\).

\[0.30 \text{ mol } C_4H_{10} \times \frac{8 \text{ mol } CO_2}{2 \text{ mol } C_4H_{10}} = 1.2 \text{ mol } CO_2\]

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 \(H_2O\).

\[0.69 \text{ mol } C_2H_5OH \times \frac{3 \text{ mol } H_2O}{1 \text{ mol } C_2H_5OH} = 2.07 = 2.1 \text{ mol } H_2O\]

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 \(O_2\).

\[3.91 \text{ mol } Fe_2O_3 \times \frac{3 \text{ mol } O_2}{2 \text{ mol } Fe_2O_3} = 5.865 = 5.87 \text{ mol } O_2\]

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 \(NiCl_2\).

\[0.479 \text{ mol } Ni_3(PO_4)_2 \times \frac{3 \text{ mol } NiCl_2}{1 \text{ mol } Ni_3(PO_4)_2} = 1.437 = 1.44 \text{ mol } NiCl_2\]
3.77 \[ 3 \text{NO}_2 + \text{H}_2\text{O} \rightarrow 2 \text{HNO}_3 + \text{NO} \]

three mol of \( \text{NO}_2 \) are equivalent to two mol of \( \text{HNO}_3 \) (from equation).

one mol of \( \text{NO}_2 \) is equivalent to 46.01 g \( \text{NO}_2 \) (from molecular weight of \( \text{NO}_2 \)).

one mol of \( \text{HNO}_3 \) is equivalent to 63.02 g \( \text{HNO}_3 \) (from molecular weight of \( \text{HNO}_3 \)).

\[
7.50 \text{ g HNO}_3 \times \frac{1 \text{ mol HNO}_3}{63.02 \text{ g HNO}_3} \times \frac{3 \text{ mol NO}_2}{2 \text{ mol HNO}_3} \times \frac{46.01 \text{ g NO}_2}{1 \text{ mol NO}_2} = 8.213
\]

\[= 8.21 \text{ g NO}_2\]

3.78 \[ 2\text{Ca}_3(\text{PO}_4)_2 + 6\text{SiO}_2 + 10\text{C} \rightarrow \text{P}_4 + 6\text{CaSiO}_3 + 10\text{CO} \]

two mol of \( \text{Ca}_3(\text{PO}_4)_2 \) are equivalent to one mol of \( \text{P}_4 \) (from equation).

one mol of \( \text{P}_4 \) is equivalent to 123.9 g \( \text{P}_4 \) (from molecular weight of \( \text{P}_4 \)).

one mol of \( \text{Ca}_3(\text{PO}_4)_2 \) is equivalent to 310.2 g \( \text{Ca}_3(\text{PO}_4)_2 \) [from molecular weight of \( \text{Ca}_3(\text{PO}_4)_2 \)].

\[
15.0 \text{ g P}_4 \times \frac{1 \text{ mol P}_4}{123.88 \text{ g P}_4} \times \frac{2 \text{ mol Ca}_3(\text{PO}_4)_2}{1 \text{ mol P}_4} \times \frac{310.18 \text{ g Ca}_3(\text{PO}_4)_2}{1 \text{ mol Ca}_3(\text{PO}_4)_2} = 75.11 = 75.1 \text{ g Ca}_3(\text{PO}_4)_2
\]

3.79 \[ \text{WO}_3 + 3\text{H}_2 \rightarrow \text{W} + 3\text{H}_2\text{O} \]

one mol of \( \text{W} \) is equivalent to three moles of \( \text{H}_2 \) (from equation).

one mol of \( \text{H}_2 \) is equivalent to 2.016 g \( \text{H}_2 \) (from molecular weight of \( \text{H}_2 \)).

one mol of \( \text{W} \) is equivalent to 183.8 g \( \text{W} \) (from atomic weight of \( \text{W} \)).

4.81 kg of \( \text{H}_2 \) is equivalent to \( 4.81 \times 10^3 \) g of \( \text{H}_2 \).

\[
4.81 \times 10^3 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \times \frac{1 \text{ mol W}}{3 \text{ mol H}_2} \times \frac{183.85 \text{ g W}}{1 \text{ mol W}} = 1.462 \times 10^5
\]

\[= 1.46 \times 10^5 \text{ g W}\]
3.80 \[ 4\text{C}_3\text{H}_6 + 6\text{NO} \rightarrow 4\text{C}_3\text{H}_3\text{N} + 6\text{H}_2\text{O} + \text{N}_2 \]

four mol of \( \text{C}_3\text{H}_6 \) are equivalent to four mol of \( \text{C}_3\text{H}_3\text{N} \) (from equation).

one mol of \( \text{C}_3\text{H}_6 \) is equivalent to 42.08 g \( \text{C}_3\text{H}_6 \) (from molecular weight of \( \text{C}_3\text{H}_6 \)).

one mol of \( \text{C}_3\text{H}_3\text{N} \) is equivalent to 53.06 g \( \text{C}_3\text{H}_3\text{N} \) (from molecular weight of \( \text{C}_3\text{H}_3\text{N} \)).

651 kg of \( \text{C}_3\text{H}_6 \) are equivalent to \( 6.51 \times 10^5 \) g \( \text{C}_3\text{H}_6 \).

\[
6.51 \times 10^5 \text{ g C}_3\text{H}_6 \times \frac{1 \text{ mol C}_3\text{H}_6}{42.08 \text{ g C}_3\text{H}_6} \times \frac{4 \text{ mol C}_3\text{H}_3\text{N}}{4 \text{ mol C}_3\text{H}_6} \times \frac{53.06 \text{ g C}_3\text{H}_3\text{N}}{1 \text{ mol C}_3\text{H}_3\text{N}}
= 8.208 \times 10^5 = 8.21 \times 10^5 \text{ g C}_3\text{H}_3\text{N}
\]

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

\[ \text{CS}_2 + 3\text{Cl}_2 \rightarrow \text{CCl}_4 + \text{S}_2\text{Cl}_2 \]

\[
62.7 \text{ g Cl}_2 \times \frac{1 \text{ mol Cl}_2}{70.90 \text{ g Cl}_2} \times \frac{1 \text{ mol CS}_2}{3 \text{ mol Cl}_2} \times \frac{76.15 \text{ g CS}_2}{1 \text{ mol CS}_2} = 22.48 = 22.4 \text{ g CS}_2
\]

3.82 From the molecular models, the balanced chemical equation is

\[ 4\text{NH}_3 + 5\text{O}_2 \xrightarrow{\text{Pt}} 4\text{NO} + 6\text{H}_2\text{O} \]

The molar mass of \( \text{NH}_3 \) is 17.03 g/mol, and for \( \text{O}_2 \), it is 32.00 g/mol. This gives

\[
6.1 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3} \times \frac{5 \text{ mol O}_2}{4 \text{ mol NH}_3} \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} = 14.32 = 14.3 \text{ g O}_2
\]

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

\[ 2\text{N}_2\text{O}_5 \rightarrow 4\text{NO}_2 + \text{O}_2 \]

\[
1.315 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{4 \text{ mol NO}_2}{1 \text{ mol O}_2} \times \frac{46.01 \text{ g NO}_2}{1 \text{ mol NO}_2} = 7.5628 = 7.563 \text{ g NO}_2
\]
Write the equation, and set up the calculation below the equation (after calculating the two formula weights):

\[ 3\text{Cu} + 8\text{HNO}_3 \rightarrow 3\text{Cu(NO}_3)_2 + 2\text{NO} + 4\text{H}_2\text{O} \]

\[
5.92 \text{ g Cu(NO}_3)_2 \times \frac{1\text{ mol Cu(NO}_3)_2}{187.56 \text{ g Cu(NO}_3)_2} \times \frac{2\text{ mol NO}}{3\text{ mol Cu(NO}_3)_2} \times \frac{30.01 \text{ g NO}}{1\text{ mol NO}} = 0.6314 = 0.631 \text{ g NO}
\]

First determine whether KO\(_2\) or H\(_2\)O is the limiting reactant by calculating the moles of \text{O}_2 that each would form if it were the limiting reactant. Identify the limiting reactant by the smaller number of moles of \text{O}_2 formed.

\[
0.15 \text{ mol H}_2\text{O} \times \frac{3\text{ mol O}_2}{2\text{ mol H}_2\text{O}} = 0.225 \text{ mol O}_2
\]

\[
0.25 \text{ mol KO}_2 \times \frac{3\text{ mol O}_2}{4\text{ mol KO}_2} = 0.187 \text{ mol O}_2 (\text{KO}_2 \text{ is the limiting reactant})
\]

The moles of \text{O}_2 produced = 0.19 mol.

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.

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

\[
1.47 \text{ mol Cl}_2 \times \frac{1\text{ mol NaClO}}{1\text{ mol Cl}_2} = 1.47 \text{ mol NaClO}
\]

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

First determine whether CO or H\(_2\) is the limiting reactant by calculating the moles of CH\(_3\)OH that each would form if it were the limiting reactant. Identify the limiting reactant by the smaller number of moles of CH\(_3\)OH formed. Use the molar mass of CH\(_3\)OH to calculate the mass of CH\(_3\)OH formed. Then, calculate the mass of the unconsumed reactant.

\[
\text{CO} + 2\text{H}_2 \rightarrow \text{CH}_3\text{OH}
\]

(continued)
\[ 10.2 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.016 \text{ g } H_2} \times \frac{1 \text{ mol } CH_3OH}{2 \text{ mol } H_2} = 2.529 \text{ mol } CH_3OH \]

\[ 35.4 \text{ g } CO \times \frac{1 \text{ mol } CO}{28.01 \text{ g } CO} \times \frac{1 \text{ mol } CH_3OH}{1 \text{ mol } CO} = 1.263 \text{ mol } CH_3OH \]

CO is the limiting reactant.

\[ \text{Mass } CH_3OH \text{ formed} = 1.263 \text{ mol } CH_3OH \times \frac{32.042 \text{ g } CH_3OH}{1 \text{ mol } CH_3OH} = 40.47 = 40.5 \text{ g } CH_3OH \]

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:

\[ 1.263 \text{ mol } CH_3OH \times \frac{2 \text{ mol } H_2}{1 \text{ mol } CH_3OH} \times \frac{2.016 \text{ g } H_2}{1 \text{ mol } H_2} = 5.092 \text{ g } H_2 \]

The unreacted H\(_2\) = 10.2 g total H\(_2\) - 5.092 g reacted H\(_2\) = 5.108 = 5.1 g H\(_2\).

\[ \text{First, determine whether } CS_2 \text{ or } O_2 \text{ is the limiting reactant by calculating the moles of } \text{SO}_2 \text{ that each would form if it were the limiting reactant. Identify the limiting reactant by the smaller number of moles of } \text{SO}_2 \text{ formed. Use the molar mass of } \text{SO}_2 \text{ to calculate the mass of } \text{SO}_2 \text{ formed. Then, calculate the mass of the unconsumed reactant.} \]

\[ CS_2 + 3O_2 \rightarrow CO_2 + 2SO_2 \]

\[ 30.0 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} \times \frac{2 \text{ mol } SO_2}{3 \text{ mol } O_2} = 0.6250 \text{ mol } SO_2 \]

\[ 35.0 \text{ g } CS_2 \times \frac{1 \text{ mol } CS_2}{76.15 \text{ g } CS_2} \times \frac{2 \text{ mol } SO_2}{1 \text{ mol } CS_2} = 0.9192 \text{ mol } SO_2 \]

\( O_2 \) is the limiting reactant.

\[ \text{Mass } SO_2 \text{ formed} = 0.6250 \text{ mol } SO_2 \times \frac{64.07 \text{ g } SO_2}{1 \text{ mol } SO_2} = 40.04 = 40.0 \text{ g } SO_2 \]

(continued)
CS\textsubscript{2} is left unconsumed at the end of the reaction. The mass of CS\textsubscript{2} that reacts can be calculated from the moles of product obtained:

\[
0.6250 \text{ mol SO}_2 \times \frac{1 \text{ mol CS}_2}{2 \text{ mol SO}_2} \times \frac{76.15 \text{ g CS}_2}{1 \text{ mol CS}_2} = 23.79 \text{ g CS}_2
\]

The unreacted CS\textsubscript{2} = 30.0 g total CS\textsubscript{2} - 23.79 g reacted CS\textsubscript{2} = 6.20 = 6.2 g CS\textsubscript{2}

3.89 First, determine which of the three reactants is the limiting reactant by calculating the moles of TiCl\textsubscript{4} that each would form if it were the limiting reactant. Identify the limiting reactant by the smallest number of moles of TiCl\textsubscript{4} formed. Use the molar mass of TiCl\textsubscript{4} to calculate the mass of TiCl\textsubscript{4} formed.

\[
3\text{TiO}_2 + 4\text{C} + 6\text{Cl}_2 \rightarrow 3\text{TiCl}_4 + 2\text{CO}_2 + 2\text{CO}
\]

\[
4.15 \text{ g TiO}_2 \times \frac{1 \text{ mol TiO}_2}{79.88 \text{ g TiO}_2} \times \frac{3 \text{ mol TiCl}_4}{3 \text{ mol TiO}_2} = 0.05195 \text{ mol TiCl}_4
\]

\[
5.67 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} \times \frac{3 \text{ mol TiCl}_4}{4 \text{ mol C}} = 0.35407 \text{ mol TiCl}_4
\]

\[
6.78 \text{ g Cl}_2 \times \frac{1 \text{ mol Cl}_2}{70.90 \text{ g Cl}_2} \times \frac{3 \text{ mol TiCl}_4}{6 \text{ mol Cl}_2} = 0.04781 \text{ mol TiCl}_4
\]

Cl\textsubscript{2} is the limiting reactant.

\[
\text{Mass TiCl}_4 \text{ formed} = 0.04781 \text{ mol TiCl}_4 \times \frac{189.68 \text{ g TiCl}_4}{1 \text{ mol TiCl}_4} = 9.068
\]

\[
= 9.07 \text{ g TiCl}_4
\]

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.

\[
2\text{NH}_3 + 3\text{O}_2 + 2\text{CH}_4 \rightarrow 2\text{HCN} + 6\text{H}_2\text{O}
\]

\[
11.5 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3} \times \frac{2 \text{ mol HCN}}{2 \text{ mol NH}_3} = 0.675 \text{ mol HCN}
\]

\[
10.5 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} \times \frac{2 \text{ mol HCN}}{2 \text{ mol CH}_4} = 0.654 \text{ mol HCN}
\]

(continued)
O₂ is the limiting reactant.

\[
\begin{align*}
\text{Mass HCN formed} & = 0.25 \times \frac{27.03 \text{ g HCN}}{1 \text{ mol HCN}} \\
& = 6.76 \text{ g HCN}
\end{align*}
\]

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.

\[
\begin{align*}
\text{C}_7\text{H}_6\text{O}_3 + \text{C}_4\text{H}_6\text{O}_3 & \rightarrow \text{C}_9\text{H}_8\text{O}_4 + \text{C}_2\text{H}_4\text{O}_2 \\
4.00 \text{ g C}_4\text{H}_6\text{O}_3 & \times \frac{1 \text{ mol C}_4\text{H}_6\text{O}_3}{102.09 \text{ g C}_4\text{H}_6\text{O}_3} \times \frac{1 \text{ mol C}_9\text{H}_8\text{O}_4}{1 \text{ mol C}_4\text{H}_6\text{O}_3} = 0.03918 \text{ mol C}_9\text{H}_8\text{O}_4 \\
2.00 \text{ g C}_7\text{H}_6\text{O}_3 & \times \frac{1 \text{ mol C}_7\text{H}_6\text{O}_3}{138.12 \text{ g C}_7\text{H}_6\text{O}_3} \times \frac{1 \text{ mol C}_9\text{H}_8\text{O}_4}{1 \text{ mol C}_7\text{H}_6\text{O}_3} = 0.01448 \text{ mol C}_9\text{H}_8\text{O}_4
\end{align*}
\]

Thus, C₇H₆O₃ is the limiting reactant. The theoretical yield of C₉H₈O₄ is

\[
0.01448 \text{ mol C}_9\text{H}_8\text{O}_4 \times \frac{180.15 \text{ g C}_9\text{H}_8\text{O}_4}{1 \text{ mol C}_9\text{H}_8\text{O}_4} = 2.609 \text{ g C}_9\text{H}_8\text{O}_4
\]

The percentage yield is

\[
\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{1.86 \text{ g}}{2.609 \text{ g}} \times 100\% = 71.29 = 71.3\%
\]

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.

\[
\begin{align*}
\text{C}_7\text{H}_6\text{O}_3 + \text{CH}_3\text{OH} & \rightarrow \text{C}_8\text{H}_8\text{O}_3 + \text{H}_2\text{O} \\
11.20 \text{ g CH}_3\text{OH} & \times \frac{1 \text{ mol CH}_3\text{OH}}{32.04 \text{ g CH}_3\text{OH}} \times \frac{1 \text{ mol C}_8\text{H}_8\text{O}_3}{1 \text{ mol CH}_3\text{OH}} = 0.3496 \text{ mol C}_8\text{H}_8\text{O}_3
\end{align*}
\]

(continued)
Thus, C₇H₆O₃ is the limiting reactant. The theoretical yield of C₈H₈O₃ is

\[
0.01086 \text{ mol C}_8\text{H}_8\text{O}_3 \times \frac{152.14 \text{ g C}_8\text{H}_8\text{O}_3}{1 \text{ mol C}_8\text{H}_8\text{O}_3} = 1.652 \text{ g C}_8\text{H}_8\text{O}_3
\]

The percentage yield is

\[
\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{1.27 \text{ g}}{1.651 \text{ g}} \times 100\% = 76.92 = 76.9\%
\]

### 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:

\[
\begin{align*}
8 \text{ mol C} & \times 12.01 \text{ g C/1 mol C} = 96.08 \text{ g C} \\
10 \text{ mol H} & \times 1.008 \text{ g H/1 mol H} = 10.08 \text{ g H} \\
4 \text{ mol N} & \times 14.01 \text{ g N/1 mol N} = 56.04 \text{ g N} \\
2 \text{ mol O} & \times 16.00 \text{ g O/1 mol O} = 32.00 \text{ g O}
\end{align*}
\]

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 ÷ molar 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.)
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:

\[
\begin{align*}
17 \text{ mol C} \times 12.01 \text{ g C/1 mol C} &= 204.17 \text{ g C} \\
19 \text{ mol H} \times 1.008 \text{ g H/1 mol H} &= 19.15 \text{ g H} \\
1 \text{ mol N} \times 14.01 \text{ g N/1 mol N} &= 14.01 \text{ g N} \\
3 \text{ mol O} \times 16.00 \text{ g O/1 mol O} &= 48.00 \text{ g O} \\
\text{1 mol of morphine (total)} &= 285.33 \text{ g (molar mass)}
\end{align*}
\]

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.)

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:

\[
\begin{align*}
49.1 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} &= 4.088 \text{ mol C} \\
2.7 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} &= 2.68 \text{ mol H} \\
48.2 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} &= 1.360 \text{ mol Cl}
\end{align*}
\]

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

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

$$= 73.496 \approx 73.50 \text{ amu}$$

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.

$$64.3 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 5.353 \text{ mol C}$$

$$7.2 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 7.14 \text{ mol H}$$

$$28.5 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.781 \text{ mol O}$$

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

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

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

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:

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

$$= 56.062 \approx 56.06 \text{ amu}$$

(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:

$$0.01665 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 0.004544 \text{ g C}$$

Percent C = (0.004544 g C ÷ 0.00796 g comp.) x 100% = 57.09%

$$0.01196 \text{ g BaSO}_4 \times \frac{1 \text{ mol BaSO}_4}{233.39 \text{ g BaSO}_4} \times \frac{1 \text{ mol S}}{1 \text{ mol BaSO}_4} \times \frac{32.07 \text{ g S}}{1 \text{ mol S}}$$

= 0.001643 g S

Percent S = (0.001643 g S ÷ 0.00431 g comp.) x 100% = 38.12%

Percent H = 100.00% - (57.09 + 38.12)% = 4.79%

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

$$57.09 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.754 \text{ mol C}$$

$$38.12 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g C}} = 1.189 \text{ mol S}$$

$$4.79 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.752 \text{ mol H}$$

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 $C_6H_8O_2$. 
3.98 Find the percentage composition of H and N from the analysis:

\[
0.00663 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.0007417 \text{ g H}
\]

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

Percent H = \( \frac{0.0007417 \text{ g H}}{0.00146 \text{ g N}} \times 100\% = 7.64\% \)

Percent N = \( \frac{0.00146 \text{ g N}}{0.00146 \text{ g N}} \times 100\% = 15.0\% \)

Percent C = 100.00\% - (7.64 + 15.0)\% = 77.4\%

Calculate the moles from grams to obtain the empirical formula:

\[
\begin{align*}
77.4 \text{ g C} & \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 6.44 \text{ mol C} \\
7.64 \text{ g H} & \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 7.58 \text{ mol H} \\
15.0 \text{ g N} & \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 1.07 \text{ mol N}
\end{align*}
\]

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 C₆H₇N (formula weight = 93 amu). Because the molecular weight was given as 93 amu, the molecular formula is also C₆H₇N.

3.99 For g CaCO₃, use this equation: \( \text{CaCO}_3 + \text{H}_2\text{C}_2\text{O}_4 \rightarrow \text{CaC}_2\text{O}_4 + \text{H}_2\text{O} + \text{CO}_2 \).

\[
0.472 \text{ g CaC}_2\text{O}_4 \times \frac{1 \text{ mol CaC}_2\text{O}_4}{128.10 \text{ g CaC}_2\text{O}_4} \times \frac{1 \text{ mol CaCO}_3}{1 \text{ mol CaC}_2\text{O}_4} \times \frac{100.09 \text{ g CaCO}_3}{1 \text{ mol CaCO}_3} = 0.3688 \text{ g CaCO}_3
\]

Mass percentage CaCO₃ = \( \frac{\text{mass CaCO}_3}{\text{mass limestone}} \times 100\% = \frac{0.3688 \text{ g}}{0.438 \text{ g}} \times 100\% = 84.19 \% = 84.2\% \)
3.100 For the mass of TiO$_2$, use this equation: TiO$_2$ + C + 2Cl$_2$ → TiCl$_4$ + CO$_2$.

$$35.4 \text{ g TiCl}_4 \times \frac{1 \text{ mol TiCl}_4}{189.68 \text{ g TiCl}_4} \times \frac{1 \text{ mol TiO}_2}{1 \text{ mol TiCl}_4} \times \frac{79.88 \text{ g TiO}_2}{1 \text{ mol TiO}_2} = 14.91 \text{ g TiO}_2$$

Mass percent TiO$_2$ = \(\frac{\text{mass TiO}_2}{\text{mass rutile}}\) x 100% = \(\frac{14.91 \text{ g}}{17.4 \text{ g}}\) x 100% = 85.68% = 85.7%

3.101 Calculate the theoretical yield using this equation: 2C$_2$H$_4$ + O$_2$ → 2C$_2$H$_4$O.

$$10.6 \text{ g C}_2\text{H}_4 \times \frac{1 \text{ mol C}_2\text{H}_4}{28.05 \text{ g C}_2\text{H}_4} \times \frac{2 \text{ mol C}_2\text{H}_4\text{O}}{2 \text{ mol C}_2\text{H}_4} \times \frac{44.05 \text{ g C}_2\text{H}_4\text{O}}{1 \text{ mol C}_2\text{H}_4\text{O}}$$

= 16.65 g C$_2$H$_4$O

Percent yield = \(\frac{\text{actual yield}}{\text{theoretical yield}}\) x 100% = \(\frac{9.91 \text{ g}}{16.65 \text{ g}}\) x 100% = 59.53 = 59.5%

3.102 Calculate the theoretical yield using this equation:

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

$$22.4 \text{ g C}_6\text{H}_6 \times \frac{1 \text{ mol C}_6\text{H}_6}{78.11 \text{ g C}_6\text{H}_6} \times \frac{1 \text{ mol C}_6\text{H}_5\text{NO}_2}{1 \text{ mol C}_6\text{H}_6} \times \frac{123.11 \text{ g C}_6\text{H}_5\text{NO}_2}{1 \text{ mol C}_6\text{H}_5\text{NO}_2}$$

= 35.30 g C$_6$H$_5$NO$_2$

Percent yield = \(\frac{\text{actual yield}}{\text{theoretical yield}}\) x 100% = \(\frac{31.6 \text{ g}}{35.30 \text{ g}}\) x 100% = 89.51 = 89.5%

3.103 To find Zn, use these equations:

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

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

$$\text{75.0 g } ZnO \times \frac{1 \text{ mol } ZnO}{81.39 \text{ g } ZnO} \times \frac{1 \text{ mol } C}{1 \text{ mol } ZnO} \times \frac{12.01 \text{ g } C}{1 \text{ mol } C} = 11.07 \text{ g } C$$

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:

$$\text{75.0 g } ZnO \times \frac{1 \text{ mol } ZnO}{81.39 \text{ g } ZnO} \times \frac{1 \text{ mol } Zn}{1 \text{ mol } ZnO} \times \frac{65.39 \text{ g } Zn}{1 \text{ mol } Zn} = 60.256 = 60.3 \text{ g } Zn$$

3.104 To find $CH_4$, use these equations:

$$4 \text{NH}_3 + 5O_2 \rightarrow 4 \text{NO} + 6 \text{H}_2\text{O}$$

$$2 \text{NO} + 2 \text{CH}_4 \rightarrow 2 \text{HCN} + 2 \text{H}_2\text{O} + \text{H}_2$$

Four mol $\text{NH}_3$ produces four mol $\text{NO}$; because two mol $\text{CH}_4$ reacts with two mol $\text{NO}$, four mol $\text{CH}_4$ will react with four mol $\text{NO}$. Thus, four mol $\text{NH}_3$ is equivalent to four mol $\text{CH}_4$. Using this to calculate the mass of $\text{CH}_4$ from the mass of $\text{NH}_3$, we have

$$24.2 \text{ g } \text{NH}_3 \times \frac{1 \text{ mol } \text{NH}_3}{17.03 \text{ g } \text{NH}_3} \times \frac{4 \text{ mol } \text{CH}_4}{4 \text{ mol } \text{NH}_3} \times \frac{16.04 \text{ g } \text{CH}_4}{1 \text{ mol } \text{CH}_4} = 22.8 \text{ g } \text{CH}_4$$

Thus, all of the $\text{NH}_3$ is used up in reacting with just 22.8 g of $\text{CH}_4$, making $\text{NH}_3$ the limiting reactant. Use the mass of $\text{NH}_3$ to calculate the mass of $\text{HCN}$ formed:

$$24.2 \text{ g } \text{NH}_3 \times \frac{1 \text{ mol } \text{NH}_3}{17.03 \text{ g } \text{NH}_3} \times \frac{2 \text{ mol } \text{CH}_4}{2 \text{ mol } \text{NH}_3} \times \frac{27.03 \text{ g } \text{HCN}}{1 \text{ mol } \text{HCN}} = 38.41 = 38.4 \text{ g } \text{HCN}$$

3.105 For $\text{CaO} + 3 \text{C} \rightarrow \text{CaC}_2 + \text{CO}$, find the limiting reactant in terms of moles of $\text{CaC}_2$ obtainable:

$$\text{Mol } \text{CaC}_2 = 2.60 \times 10^3 \text{ g } \text{C} \times \frac{1 \text{ mol } \text{C}}{12.01 \text{ g } \text{C}} \times \frac{1 \text{ mol } \text{CaC}_2}{3 \text{ mol } \text{C}} = 72.16 \text{ mol}$$

$$\text{Mol } \text{CaC}_2 = 2.60 \times 10^3 \text{ g } \text{CaO} \times \frac{1 \text{ mol } \text{CaO}}{56.08 \text{ g } \text{CaO}} \times \frac{1 \text{ mol } \text{CaC}_2}{1 \text{ mol } \text{CaO}} = 46.362 \text{ mol}$$

(continued)
Because CaO is the limiting reactant, calculate the mass of CaC\textsubscript{2} from it:

\[
\text{Mass CaC}_2 = 46.362 \text{ mol CaC}_2 \times \frac{64.10 \text{ g CaC}_2}{1 \text{ mol CaC}_2} = 2.971 \times 10^3 \\
= 2.97 \times 10^3 \text{ g CaC}_2
\]

3.106 For CaF\textsubscript{2} + H\textsubscript{2}SO\textsubscript{4} \rightarrow 2HF + CaSO\textsubscript{4}, find the limiting reactant in terms of moles of HF obtainable:

\[
\text{Mol HF} = 12.8 \text{ g CaF}_2 \times \frac{1 \text{ mol CaF}_2}{78.08 \text{ g CaF}_2} \times \frac{2 \text{ mol HF}}{1 \text{ mol CaF}_2} = 0.3279 \text{ mol}
\]

\[
\text{Mol HF} = 13.2 \text{ g H}_2\text{SO}_4 \times \frac{1 \text{ mol H}_2\text{SO}_4}{98.09 \text{ g H}_2\text{SO}_4} \times \frac{2 \text{ mol HF}}{1 \text{ mol H}_2\text{SO}_4} = 0.26914 \text{ mol}
\]

Because H\textsubscript{2}SO\textsubscript{4} is the limiting reactant, calculate the mass of HF from it:

\[
\text{Mass HF} = 0.26914 \text{ mol HF} \times \frac{20.01 \text{ g HF}}{1 \text{ mol HF}} = 5.3854 = 5.39 \text{ g HF}
\]

3.107 From the equation 2Na + H\textsubscript{2}O \rightarrow 2NaOH + H\textsubscript{2}, convert the mass of H\textsubscript{2} to mass of Na, and then use the mass to calculate the percentage:

\[
0.108 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \times \frac{2 \text{ mol Na}}{1 \text{ mol H}_2} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 2.463 \text{ g Na}
\]

\[
\text{Percent Na} = \frac{\text{mass Na}}{\text{mass amalgam}} \times 100\% = \frac{2.463 \text{ g}}{15.23 \text{ g}} \times 100\% = 16.17 = 16.2\%
\]

3.108 From the equation CaCO\textsubscript{3} \rightarrow CaO + CO\textsubscript{2}, convert the mass of CO\textsubscript{2} to mass of CaCO\textsubscript{3}, and then use the mass to calculate the percentage:

\[
0.00395 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol CaCO}_3}{1 \text{ mol CO}_2} \times \frac{100.09 \text{ g CaCO}_3}{1 \text{ mol CaCO}_3} = 0.008983 \text{ g CaCO}_3
\]

(continued)
CHAPTER 3

Percent CaCO$_3$ = \( \frac{\text{mass CaCO}_3}{\text{mass sandstone}} \times 100\% = \frac{0.008983 \text{ g}}{0.0187 \text{ g}} \times 100\% \)

\[ = 48.039 = 48.0\% \]

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$_2$:

Percent SiO$_2$(silica) = 100.00\% - 48.039\% = 51.96 = 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$_2$O 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:

\[ 0.425 = y \left[ \frac{63.55 \text{ g Cu}}{79.55 \text{ g CuO}} \right] + (0.500 - y) \left[ \frac{127.10 \text{ g Cu}}{143.10 \text{ g Cu}_2\text{O}} \right] \]

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

\[ 0.425 = 0.79887 y + 0.888190 (0.500 - y) \]
\[ 0.08932 y = 0.019095 \]
\[ y = 0.2139 = 0.21 \text{ g} = \text{mass of CuO} \]

3.110 Let y equal the mass of Fe$_2$O$_3$ 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 \times 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:

\[ 0.360 = y \left[ \frac{111.70 \text{ g Fe}}{159.70 \text{ g Fe}_2\text{O}_3} \right] + (0.500 - y) \left[ \frac{55.85 \text{ g Fe}}{71.85 \text{ g FeO}} \right] \]

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

\[ 0.360 = 0.6994 y + 0.7773 (0.500 - y) \]
\[ 0.07790 y = 0.02865 \]
\[ y = 0.3677 = 0.368 \text{ g} = \text{mass of Fe}_2\text{O}_3 \]
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):

$$3.19 \times 10^{-3} \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 5.712 \times 10^{-5} \text{ mol Fe or heme}$$

Molar mass of heme = \(\frac{35.2 \times 10^{-3} \text{ g}}{5.712 \times 10^{-5} \text{ mol}} = 616.2 = 616 \text{ g/mol}\)

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:

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

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

Mass percentage S = \(\frac{0.7502 \times 10^{-3} \text{ g S}}{8.19 \times 10^{-3} \text{ g pen. V}} \times 100\% = 9.160 = 9.16\%\)

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:

$$0.7502 \times 10^{-3} \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = 2.339 \times 10^{-5} \text{ mol S}$$

Molar mass pen. V = \(\frac{8.19 \times 10^{-3} \text{ g}}{2.339 \times 10^{-5} \text{ mol}} = 350.1 = 350. \text{ g/mol}\)

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\(_2\) = \(\frac{1.11 \text{ g}}{0.0158 \text{ mol}} = 70.25 \text{ g/mol}\)

Molar mass X = \(\frac{70.25}{2} = 35.14 = 35.1 \text{ g/mol}\)

(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 ÷ 0.0316 mol = 107.9 = 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 ÷ 2 = 0.0790 mol

Molar mass of M$^{2+}$ = 1.92 g ÷ 0.0790 mol = 24.30 = 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$_2$ = 14.545 g

Mass of X in MX$_2$ = 14.545 g - 1.92 g = 12.625 g

The moles of X in MX$_2$ = 0.158 mol

Molar mass of X = 12.625 g ÷ 0.158 mol = 79.905 = 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 x 10$^3$ cm$^3$

Mass Fe = 3.00 x 10$^3$ cm$^3$ x $\frac{8.17 \text{ g alloy}}{1 \text{ cm}^3}$ x $\frac{54.7 \text{ g Fe}}{100.0 \text{ g alloy}}$ = 1.3407 x 10$^4$ g

No. of Fe atoms = 1.3407 x 10$^4$ g Fe x $\frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}}$ x $\frac{6.022 \times 10^{23} \text{ Fe atoms}}{1 \text{ mol Fe}}$

= 1.4456 x 10$^{26}$ = 1.45 x 10$^{26}$ Fe atoms
Calculations with Chemical Formulas and Equations

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.

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

\[
\text{Mass Co} = 1.9635 \text{ cm}^3 \times \frac{8.20 \text{ g alloy}}{1 \text{ cm}^3} \times \frac{12.0 \text{ g Co}}{100.0 \text{ g alloy}} = 1.932 \times 10^2 \text{ g}
\]

\[
\text{No. of Co atoms} = 1.932 \times 10^2 \text{ g Co} \times \frac{1 \text{ mol Co}}{58.93 \text{ g Co}} \times \frac{6.022 \times 10^{23} \text{ Co atoms}}{1 \text{ mol Co}}
\]

\[
= 1.974 \times 10^{24} = 1.97 \times 10^{24} \text{ Co atoms}
\]