WORKED EXAMPLE 6.1 Balancing a Chemical Equation

Propane, C_3H_8, is a colorless, odorless gas often used as a heating and cooking fuel in campers and rural homes. Write a balanced equation for the combustion reaction of propane with oxygen to yield carbon dioxide and water.

Strategy and Solution

Follow the four steps described in the text:

Step 1 Write the unbalanced equation using correct chemical formulas for all substances:

$$C_3H_8 + O_2 \longrightarrow CO_2 + H_2O \quad \text{Unbalanced}$$

Step 2 Find coefficients to balance the equation. It's usually best to begin with the most complex substance—in this case C_3H_8—and to deal with one element at a time. Look at the unbalanced equation, and note that there are 3 carbon atoms on the left side of the equation but only 1 on the right side. If we add a coefficient of 3 to CO_2 on the right, the carbons balance:

$$C_3H_8 + O_2 \longrightarrow 3\,CO_2 + H_2O \quad \text{Balanced for C}$$

Next, look at the number of hydrogen atoms. There are 8 hydrogens on the left but only 2 (in H_2O) on the right. By adding a coefficient of 4 to the H_2O on the right, the hydrogens balance:

$$C_3H_8 + O_2 \longrightarrow 3\,CO_2 + 4\,H_2O \quad \text{Balanced for C and H}$$

Finally, look at the number of oxygen atoms. There are 2 on the left but 10 on the right. By adding a coefficient of 5 to the O2 on the left, the oxygens balance:

$$C_3H_8 + 5\,O_2 \longrightarrow 3\,CO_2 + 4\,H_2O \quad \text{Balanced for C, H, and O}$$

Step 3 Make sure the coefficients are reduced to their smallest whole-number values. In fact, our answer is already correct, but we might have arrived at a different answer through trial and error:

$$2\,C_3H_8 + 10\,O_2 \longrightarrow 6\,CO_2 + 8\,H_2O$$

Although the preceding equation is balanced, the coefficients are not the smallest whole numbers. It would be necessary to divide all coefficients by 2 to reach the final equation.

Step 4 Check your answer. Count the numbers and kinds of atoms on both sides of the equation to make sure they're the same:

$$\underbrace{C_3H_8 + 5\,O_2}_{\text{3 C, 8 H, and 10 O atoms on this side}} \longrightarrow \underbrace{3\,CO_2 + 4\,H_2O}_{\text{3 C, 8 H, and 10 O atoms on this side}}$$

WORKED EXAMPLE 6.2 Balancing a Chemical Equation

The major ingredient in ordinary safety matches is potassium chlorate, $KClO_3$, a substance that can act as a source of oxygen in combustion reactions. Its reaction with ordinary table sugar (sucrose, $C_{12}H_{22}O_{11}$), for example, occurs violently to yield potassium chloride, carbon dioxide, and water. Write a balanced equation for the reaction.

Strategy and Solution

Step 1 Write the unbalanced equation, making sure the formulas for all substances are correct:

$$KClO_3 + C_{12}H_{22}O_{11} \longrightarrow KCl + CO_2 + H_2O \quad \text{Unbalanced}$$

Step 2 Find coefficients to balance the equation by starting with the most complex substance (sucrose) and considering one element at a time. Since there are 12 C atoms on the left and only 1 on the right, we can balance for carbon by adding a coefficient of 12 to CO_2 on the right:

$$KClO_3 + C_{12}H_{22}O_{11} \longrightarrow KCl + 12\ CO_2 + H_2O \quad \text{Balanced for C}$$

Since there are 22 H atoms on the left and only 2 on the right, we can balance for hydrogen by adding a coefficient of 11 to H_2O on the right:

$$KClO_3 + C_{12}H_{22}O_{11} \longrightarrow KCl + 12\ CO_2 + 11\ H_2O \quad \text{Balanced for C and H}$$

There are now 35 O atoms on the right but only 14 on the left (11 in sucrose and 3 in $KClO_3$). Thus, 21 oxygens must be added on the left. We can do this without disturbing the C and H balance by adding 7 more $KClO_3$, giving a coefficient of 8 for $KClO_3$ on the left:

$$8\ KClO_3 + C_{12}H_{22}O_{11} \longrightarrow KCl + 12\ CO_2 + 11\ H_2O \quad \text{Balanced for C, H and O}$$

Potassium and chlorine can both be balanced by adding a coefficient of 8 to KCl on the right:

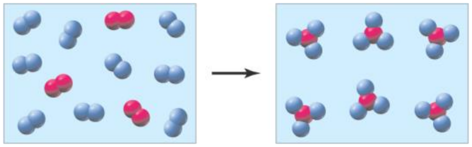
$$8\ KClO_3 + C_{12}H_{22}O_{11} \longrightarrow 8\ KCl + 12\ CO_2 + 11\ H_2O \quad \text{Balanced for C, H, O, K, and Cl}$$

Steps 3 and 4 The coefficients in the balanced equation are already reduced to their smallest whole-number values, and a check shows that the numbers and kinds of atoms are the same on both sides of the equation.

$$\underbrace{8\ KClO_3 + C_{12}H_{22}O_{11}}_{\text{8 K, 8 Cl, 12 C, 22 H, and 35 O atoms}} \longrightarrow \underbrace{8\ KCl + 12\ CO_2 + 11\ H_2O}_{\text{8 K, 8 Cl, 12 C, 22 H, and 35 O atoms}}$$

WORKED KEY CONCEPT EXAMPLE 6.3 Balancing a Chemical Equation

Write a balanced equation for the reaction of element A (red spheres) with element B (blue spheres) as represented below:



Strategy

Balancing the reactions shown in graphic representations of this sort is just a matter of counting the numbers of reactant and product units. In this example, the reactant box contains three red A_2 molecules and nine blue B_2 molecules, while the product box contains six AB_3 molecules with no reactant left over.

Solution

$$3\,A_2 + 9\,B_2 \longrightarrow 6\,AB_3 \quad \text{or} \quad A_2 + 3\,B_2 \longrightarrow 2\,AB_3$$

PROBLEM 6.1 Sodium chlorate, $NaClO_3$, decomposes when heated to yield sodium chloride and oxygen, a reaction used to provide oxygen for the emergency breathing masks in many airliners. Balance the equation.

WORKED KEY CONCEPT EXAMPLE 6.3 Balancing a Chemical Equation

Continued

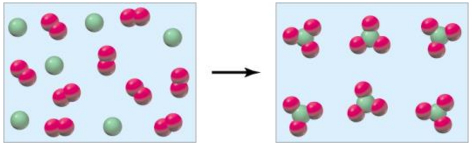
PROBLEM 6.2 Balance the following equations:

(a) $C_6H_{12}O_6 \rightarrow C_2H_6O + CO_2$ (fermentation of sugar to yield ethyl alcohol)

(b) $CO_2 + H_2O \rightarrow C_6H_{12}O_6 + O_2$ (photosynthesis reaction in green plants)

(c) $NH_3 + Cl_2 \rightarrow N_2H_4 + NH_4Cl$ (synthesis of hydrazine for rocket fuel)

KEY CONCEPT PROBLEM 6.3 Write a balanced equation for the reaction of element A (red spheres) with element B (green spheres) as represented below:



WORKED EXAMPLE 6.4 Calculating a Molecular Mass

What is the molecular mass of table sugar (sucrose, $C_{12}H_{22}O_{11}$), and what is its molar mass in g/mol?

Strategy

The molecular mass of a substance is the sum of the atomic masses of the constituent atoms. List the elements present in the molecule, and look up the atomic mass of each (we'll round off to one decimal place for convenience):

$$\text{C (12.0 amu); H (1.0 amu); O (16.0 amu)}$$

Then, multiply the atomic mass of each element by the number of times that element appears in the chemical formula, and total the results.

Solution

$$\begin{array}{rlcr}
C_{12} & (12 \times 12.0\text{ amu}) & = & 144.0\text{ amu} \\
H_{22} & (22 \times 1.0\text{ amu}) & = & 22.0\text{ amu} \\
O_{11} & (11 \times 16.0\text{ amu}) & = & 176.0\text{ amu} \\
\hline
\multicolumn{2}{r}{\text{Molec. mass of } C_{12}H_{22}O_{11}} & = & 342.0\text{ amu}
\end{array}$$

Because one molecule of sucrose has a mass of 342.0 amu, 1 mol of sucrose has a mass of 342.0 g. Thus, the molar mass of sucrose is 342.0 g/mol.

WORKED EXAMPLE 6.5 Converting Mass to Moles

How many moles of sucrose are in a tablespoon of sugar containing 2.85 g? (The molar mass of sucrose, $C_{12}H_{22}O_{11}$, was calculated in Worked Example 6.4.)

Strategy

The problem gives the mass of sucrose and asks for a mass-to-mole conversion. Use the molar mass of sucrose as a conversion factor, and set up an equation so that the unwanted unit cancels.

Solution

$$2.85\ \cancel{\text{g sucrose}} \times \frac{1\ \text{mol sucrose}}{342.0\ \cancel{\text{g sucrose}}} = 0.008\ 33\ \text{mol sucrose}$$
$$= 8.33 \times 10^{-3}\ \text{mol sucrose}$$

Ballpark Check

Because the molecular mass of sucrose is 342.0 amu, 1 mol of sucrose has a mass of 342.0 g. Thus, 2.85 g of sucrose is a bit less than one-hundredth of a mole, or 0.01 mol. The estimate agrees with the detailed solution.

WORKED EXAMPLE 6.6 Converting Moles to Mass

How many grams are in 0.0626 mol of $NaHCO_3$, the main ingredient in Alka-Seltzer tablets?

Strategy

The problem gives the number of moles of $NaHCO_3$ and asks for a mole-to-mass conversion. First, calculate the molar mass of $NaHCO_3$. Then use molar mass as a conversion factor, and set up an equation so that the unwanted unit cancels.

Solution

$$\text{Form. mass of NaHCO}_3 = 23.0\text{ amu} + 1.0\text{ amu} + 12.0\text{ amu} + (3 \times 16.0\text{ amu})$$

$$= 84.0\text{ amu}$$

$$\text{Molar mass of NaHCO}_3 = 84.0\text{ g/mol}$$

$$0.0626\ \cancel{\text{mol NaHCO}_3} \times \frac{84.0\text{ g NaHCO}_3}{1\ \cancel{\text{mol NaHCO}_3}} = 5.26\text{ g NaHCO}_3$$

WORKED EXAMPLE 6.7 Finding the Mass of One Reactant, Given the Mass of Another

Aqueous solutions of sodium hypochlorite (NaOCl), best known as household bleach, are prepared by reaction of sodium hydroxide with chlorine. How many grams of NaOH are needed to react with 25.0 g of Cl_2?

$$2\,NaOH(aq) + Cl_2(g) \longrightarrow NaOCl(aq) + NaCl(aq) + H_2O(l)$$

Strategy

Finding the relationship between numbers of reactant formula units always requires working in moles, using the general strategy outlined in Figure 6.1.

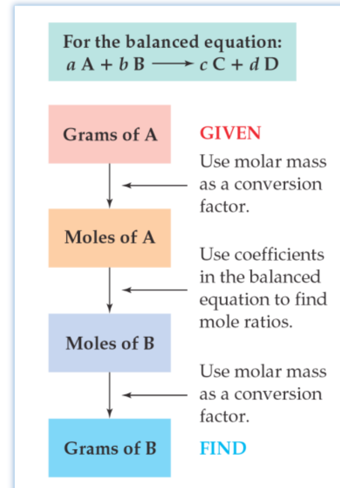


FIGURE 6.1
A summary of conversions between moles and grams for a chemical reaction. The numbers of moles tell how many molecules of each reactant are needed, as given by the coefficients of the balanced equation; the numbers of grams tell what mass of each reactant is needed.

Solution

First, find out how many moles of Cl_2 are in 25.0 g of Cl_2. This gram-to-mole conversion is done in the usual way, using the molar mass of Cl_2 (70.9 g/mol) as the conversion factor:

$$25.0\ \cancel{g\ Cl_2} \times \frac{1\ mol\ Cl_2}{70.9\ \cancel{g\ Cl_2}} = 0.353\ mol\ Cl_2$$

Next, look at the coefficients in the balanced equation. Each mole of Cl_2 reacts with 2 mol of NaOH, so 0.353 mol of Cl_2 reacts with $2 \times 0.353 = 0.706$ mol of NaOH. With the number of moles of NaOH known, carry out a mole-to-gram conversion using the molar mass of NaOH (40.0 g/mol) as a conversion factor to find that 28.2 g of NaOH is required for the reaction:

$$\text{Grams of NaOH} = 0.353\ \cancel{mol\ Cl_2} \times \frac{2\ \cancel{mol\ NaOH}}{1\ \cancel{mol\ Cl_2}} \times \frac{40.0\ g\ NaOH}{1\ \cancel{mol\ NaOH}}$$
$$= 28.2\ g\ NaOH$$

WORKED EXAMPLE 6.7 Finding the Mass of One Reactant, Given the Mass of Another

Continued

The problem can also be worked by combining the steps and setting up one large equation:

$$\text{Grams of NaOH} = 25.0\ \cancel{\text{g Cl}_2} \times \frac{1\ \cancel{\text{mol Cl}_2}}{70.9\ \cancel{\text{g Cl}_2}} \times \frac{2\ \cancel{\text{mol NaOH}}}{1\ \cancel{\text{mol Cl}_2}} \times \frac{40.0\ \text{g NaOH}}{1\ \cancel{\text{mol NaOH}}}$$
$$= 28.2\ \text{g NaOH}$$

Ballpark Check

The molar mass of NaOH is about half that of Cl_2, and 2 mol of NaOH is needed per 1 mol of Cl_2. Thus, the needed mass of NaOH will be similar to that of Cl_2, or about 25 g.

PROBLEM 6.4 Calculate the formula mass or molecular mass of the following substances:

(a) Fe_2O_3 (rust) **(b)** H_2SO_4 (sulfuric acid)
(c) $C_6H_8O_7$ (citric acid) **(d)** $C_{16}H_{18}N_2O_4S$ (penicillin G)

KEY CONCEPT PROBLEM 6.5 Aspirin can be represented by the adjacent ball-and-stick molecular model. Give the formula for aspirin, and calculate its molecular mass (red = O, gray = C, ivory = H). How many moles of aspirin are in a tablet weighing 500 mg? How many molecules?

PROBLEM 6.6 Aspirin is prepared by reaction of salicylic acid ($C_7H_6O_3$) with acetic anhydride ($C_4H_6O_3$) according to the following equation:

$$\underset{\text{Salicylic acid}}{C_7H_6O_3} + \underset{\text{Acetic anhydride}}{C_4H_6O_3} \longrightarrow \underset{\text{Aspirin}}{C_9H_8O_4} + \underset{\text{Acetic acid}}{CH_3COOH}$$

How many grams of acetic anhydride are needed to react with 4.50 g of salicylic acid? How many grams of aspirin will result? How many grams of acetic acid are formed as a by-product?

WORKED EXAMPLE 6.8 Calculating a Percent Yield

Methyl *tert*-butyl ether (MTBE, $C_5H_{12}O$), a gasoline additive now being phased out because of health concerns, can be made by reaction of isobutylene (C_4H_8) with methanol (CH_4O). What is the percent yield of the reaction if 32.8 g of methyl *tert*-butyl ether is obtained from reaction of 26.3 g of isobutylene with sufficient methanol?

$$\underset{\text{Isobutylene}}{C_4H_8(g)} + CH_4O(l) \longrightarrow \underset{\text{Methyl } \textit{tert}\text{-butyl ether (MTBE)}}{C_5H_{12}O(l)}$$

Strategy

We need to calculate the amount of methyl *tert*-butyl ether that could theoretically be produced from 26.3 g of isobutylene and compare that theoretical amount to the actual amount (32.8 g). As always, stoichiometry problems begin by calculating the molar masses of reactants and products. Coefficients of the balanced equation then tell mole ratios, and molar masses act as conversion factors between moles and masses.

Solution

Isobutylene, C_4H_8: Molec. mass $= (4 \times 12.0\text{ amu}) + (8 \times 1.0\text{ amu}) = 56.0\text{ amu}$

Molar mass of isobutylene $= 56.0\text{ g/mol}$

MTBE, $C_5H_{12}O$: Molec. mass $= (5 \times 12.0\text{ amu}) + (12 \times 1.0\text{ amu}) + 16.0\text{ amu} = 88.0\text{ amu}$

Molar mass of MTBE $= 88.0\text{ g/mol}$

To calculate the amount of MTBE that could theoretically be produced from 26.3 g of isobutylene, we first have to find the number of moles of reactant, using molar mass as the conversion factor:

$$26.3\ \cancel{\text{g isobutylene}} \times \frac{1\text{ mol isobutylene}}{56.0\ \cancel{\text{g isobutylene}}} = 0.470\text{ mol isobutylene}$$

According to the balanced equation, 1 mol of product is produced per mol of reactant, so we know that 0.470 mol of isobutylene can theoretically yield 0.470 mol of MTBE. Finding the mass of this MTBE requires a mole-to-mass conversion:

$$0.470\ \cancel{\text{mol isobutylene}} \times \frac{1\ \cancel{\text{mol MTBE}}}{1\ \cancel{\text{mol isobutylene}}} \times \frac{88.0\text{ g MTBE}}{1\ \cancel{\text{mol MTBE}}} = 41.4\text{ g MTBE}$$

Dividing the actual amount by the theoretical amount and multiplying by 100% gives the percent yield:

$$\frac{32.8\text{ g MTBE}}{41.4\text{ g MTBE}} \times 100\% = 79.2\%$$

WORKED EXAMPLE 6.9 Calculating a Yield in Grams, Given a Percent Yield

Diethyl ether ($C_4H_{10}O$), the "ether" used medically as an anesthetic, is prepared commercially by treatment of ethyl alcohol (C_2H_6O) with an acid. How many grams of diethyl ether would you obtain from 40.0 g of ethyl alcohol if the percent yield of the reaction is 87%?

$$\underset{\text{Ethyl alcohol}}{2\,C_2H_6O(l)} \xrightarrow{\text{Acid}} \underset{\text{Diethyl ether}}{C_4H_{10}O(l)} + H_2O(l)$$

Strategy

Treat this as a typical stoichiometry problem to find the amount of diethyl ether that can theoretically be formed from 40.0 g of ethyl alcohol, and then multiply the answer by 87% to find the amount actually formed.

Solution

First, calculate the molar masses of the reactant and product:

$$\begin{aligned}\text{Ethyl alcohol, } C_2H_6O\text{: Molec. mass} &= (2 \times 12.0\text{ amu}) + (6 \times 1.0\text{ amu}) + 16.0\text{ amu} \\ &= 46.0\text{ amu}\end{aligned}$$

$$\text{Molar mass of ethyl alcohol} = 46.0\text{ g/mol}$$

$$\begin{aligned}\text{Diethyl ether, } C_4H_{10}O\text{: Molec. mass} &= (4 \times 12.0\text{ amu}) + (10 \times 1.0\text{ amu}) + 16.0\text{ amu} \\ &= 74.0\text{ amu}\end{aligned}$$

$$\text{Molar mass of diethyl ether} = 74.0\text{ g/mol}$$

Next, find how many moles of ethyl alcohol are in 40.0 g by using molar mass as a conversion factor:

$$40.0\ \cancel{\text{g ethyl alcohol}} \times \frac{1\text{ mol ethyl alcohol}}{46.0\ \cancel{\text{g ethyl alcohol}}} = 0.870\text{ mol ethyl alcohol}$$

Because we started with 0.870 mol of ethyl alcohol, and because the balanced equation indicates that 2 mol of ethyl alcohol yield 1 mol of diethyl ether, we can theoretically obtain 0.435 mol of product:

$$0.870\ \cancel{\text{mol ethyl alcohol}} \times \frac{1\text{ mol diethyl ether}}{2\ \cancel{\text{mol ethyl alcohol}}} = 0.435\text{ mol diethyl ether}$$

WORKED EXAMPLE 6.9 Calculating a Yield in Grams, Given a Percent Yield

Continued

We therefore need to find how many grams of diethyl ether are in 0.435 mol, using molar mass as the conversion factor:

$$0.435\ \cancel{\text{mol diethyl ether}} \times \frac{74.0\text{ g diethyl ether}}{1\ \cancel{\text{mol diethyl ether}}} = 32.2\text{ g diethyl ether}$$

Finally, we have to multiply the theoretical amount of product by the observed yield (87% = 0.87) to find how much diethyl ether is actually formed:

$$32.2\text{ g diethyl ether} \times 0.87 = 28\text{ g diethyl ether}$$

PROBLEM 6.7 Ethyl alcohol is prepared industrially by the reaction of ethylene, C_2H_4, with water. What is the percent yield of the reaction if 4.6 g of ethylene gives 4.7 g of ethyl alcohol?

$$\underset{\text{Ethylene}}{C_2H_4(g)} + H_2O(l) \longrightarrow \underset{\text{Ethyl alcohol}}{C_2H_6O(l)}$$

PROBLEM 6.8 Dichloromethane (CH_2Cl_2), used as a solvent in the decaffeination of coffee beans, is prepared by reaction of methane (CH_4) with chlorine. How many grams of dichloromethane result from reaction of 1.85 kg of methane if the yield is 43.1%?

$$\underset{\text{Methane}}{CH_4(g)} + \underset{\text{Chlorine}}{2\,Cl_2(g)} \longrightarrow \underset{\text{Dichloromethane}}{CH_2Cl_2(l)} + 2\,HCl(g)$$

WORKED EXAMPLE 6.10 Calculating the Amount of an Excess Reactant

Cisplatin, an anticancer agent used for the treatment of solid tumors, is prepared by the reaction of ammonia with potassium tetrachloroplatinate. Assume that 10.0 g of K_2PtCl_4 and 10.0 g of NH_3 are allowed to react.

$$\underset{\text{Potassium tetrachloroplatinate}}{K_2PtCl_4(aq)} + 2\,NH_3(aq) \longrightarrow \underset{\text{Cisplatin}}{Pt(NH_3)_2Cl_2(s)} + 2\,KCl(aq)$$

(a) Which reactant is limiting, and which is in excess?
(b) How many grams of the excess reactant are consumed, and how many grams remain?
(c) How many grams of cisplatin are formed?

Strategy

When solving a problem that deals with limiting reactants, the idea is to find how many moles of all reactants are actually present and compare the mole ratios of those actual amounts to the mole ratios required by the balanced equation. That comparison will identify the reactant there is too much of (the excess reactant) and the reactant there is too little of (the limiting reactant).

Solution

(a) Finding the molar amounts of reactants always begins by calculating formula masses and using molar masses as conversion factors:

$$\text{Form. mass of } K_2PtCl_4 = (2 \times 39.1\text{ amu}) + 195.1\text{ amu} + (4 \times 35.5\text{ amu}) = 415.3\text{ amu}$$

$$\text{Molar mass of } K_2PtCl_4 = 415.3\text{ g/mol}$$

$$\text{Moles of } K_2PtCl_4 = 10.0\ \cancel{\text{g } K_2PtCl_4} \times \frac{1\text{ mol } K_2PtCl_4}{415.3\ \cancel{\text{g } K_2PtCl_4}} = 0.0241\text{ mol } K_2PtCl_4$$

$$\text{Molec. mass of } NH_3 = 14.0\text{ amu} + (3 \times 1.0\text{ amu}) = 17.0\text{ amu}$$

$$\text{Molar mass of } NH_3 = 17.0\text{ g/mol}$$

$$\text{Moles of } NH_3 = 10.0\ \cancel{\text{g } NH_3} \times \frac{1\text{ mol } NH_3}{17.0\ \cancel{\text{g } NH_3}} = 0.588\text{ mol } NH_3$$

WORKED EXAMPLE 6.10 Calculating the Amount of an Excess Reactant

Continued

These calculations tell us that we have 0.588 mol of ammonia and 0.0241 mol of K_2PtCl_4, or $0.588/0.0241 = 24.4$ times as much ammonia as K_2PtCl_4. The coefficients in the balanced equation, however, say that only two times as much ammonia as K_2PtCl_4 is needed. Thus, a large excess of NH_3 is present, and K_2PtCl_4 is the limiting reactant.

(b) With the identities of the excess reactant and limiting reactant known, we now have to first find how many moles of each undergo reaction and then carry out mole-to-gram conversions to find the mass of each reactant consumed. The entire amount of the limiting reactant (K_2PtCl_4) is used up, but only the amount of the excess reactant (NH_3) required by stoichiometry undergoes reaction:

$$\text{Moles of } K_2PtCl_4 \text{ consumed} = 0.0241 \text{ mol } K_2PtCl_4$$

$$\text{Moles of } NH_3 \text{ consumed} = 0.0241 \cancel{\text{mol } K_2PtCl_4} \times \frac{2 \text{ mol } NH_3}{1 \cancel{\text{mol } K_2PtCl_4}} = 0.0482 \text{ mol } NH_3$$

$$\text{Grams of } NH_3 \text{ consumed} = 0.0482 \cancel{\text{mol } NH_3} \times \frac{17.0 \text{ g } NH_3}{1 \cancel{\text{mol } NH_3}} = 0.819 \text{ g } NH_3$$

$$\text{Grams of } NH_3 \text{ not consumed} = (10.0 \text{ g} - 0.819 \text{ g})\ NH_3 = 9.2 \text{ g } NH_3$$

(c) The balanced equation shows that 1 mol of cisplatin is formed for each mole of K_2PtCl_4 consumed. Thus, 0.0241 mol of cisplatin is formed from 0.0241 mol of K_2PtCl_4. To determine the mass of cisplatin produced, we must calculate its molar mass and then carry out a mole-to-gram conversion:

$$\text{Molec. mass of } Pt(NH_3)_2Cl_2 = 195.1 \text{ amu} + (2 \times 17.0 \text{ amu}) + (2 \times 35.5 \text{ amu}) = 300.1 \text{ amu}$$

$$\text{Molar mass of } Pt(NH_3)_2Cl_2 = 300.1 \text{ g/mol}$$

$$\text{Grams of } Pt(NH_3)_2Cl_2 = 0.0241 \cancel{\text{mol } Pt(NH_3)_2Cl_2} \times \frac{300.1 \text{ g } Pt(NH_3)_2Cl_2}{1 \cancel{\text{mol } Pt(NH_3)_2Cl_2}} = 7.23 \text{ g } Pt(NH_3)_2Cl_2$$

WORKED EXAMPLE 6.10 Calculating the Amount of an Excess Reactant

Continued

PROBLEM 6.9 Lithium oxide is used aboard the space shuttle to remove water from the air supply according to the equation

$$Li_2O(s) + H_2O(g) \longrightarrow 2\ LiOH(s)$$

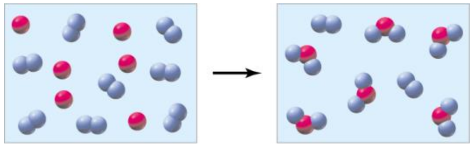
If 80.0 kg of water is to be removed and 65 kg of Li_2O is available, which reactant is limiting? How many kilograms of the excess reactant remain?

PROBLEM 6.10 After lithium hydroxide is produced aboard the space shuttle by reaction of Li_2O with H_2O (Problem 6.9), it is used to remove exhaled carbon dioxide from the air supply according to the equation

$$LiOH(s) + CO_2(g) \longrightarrow LiHCO_3(s)$$

How many grams of CO_2 can 500.0 g of LiOH absorb?

KEY CONCEPT PROBLEM 6.11 The following diagram represents the reaction of A (red spheres) with B_2 (blue spheres):



(a) Write a balanced equation for the reaction, and identify the limiting reactant.
(b) How many moles of product can be made from 1.0 mol of A and 1.0 mol of B_2?

WORKED EXAMPLE 6.11 Calculating the Molarity of a Solution

What is the molarity of a solution made by dissolving 2.355 g of sulfuric acid (H_2SO_4) in water and diluting to a final volume of 50.0 mL?

Strategy

Molarity is the number of moles of solute per liter of solution. Thus, it's necessary to find the number of moles of sulfuric acid in 2.355 g and then divide by the volume of the solution.

Solution

$$\text{Molec. mass of } H_2SO_4 = (2 \times 1.0\text{ amu}) + 32.1\text{ amu} + (4 \times 16.0\text{ amu}) = 98.1\text{ amu}$$

$$\text{Molar mass of } H_2SO_4 = 98.1\text{ g/mol}$$

$$2.355\ \cancel{\text{g } H_2SO_4} \times \frac{1\text{ mol } H_2SO_4}{98.1\ \cancel{\text{g } H_2SO_4}} = 0.0240\text{ mol } H_2SO_4$$

$$\frac{0.0240\text{ mol } H_2SO_4}{0.0500\text{ L}} = 0.480\text{ M}$$

The solution has a sulfuric acid concentration of 0.480 M.

WORKED EXAMPLE 6.12 Calculating the Number of Moles of Solute in a Solution

Hydrochloric acid is sold commercially as a 12.0 M aqueous solution. How many moles of HCl are in 300.0 mL of 12.0 M solution?

Strategy

The number of moles of solute is calculated by multiplying the molarity of the solution by its volume.

Solution

$$\text{Moles of HCl} = (\text{Molarity of solution}) \times (\text{Volume of solution})$$

$$= \frac{12.0 \text{ mol HCl}}{1 \cancel{\text{L}}} \times 0.3000 \cancel{\text{L}} = 3.60 \text{ mol HCl}$$

There are 3.60 mol of HCl in 300.0 mL of 12.0 M solution.

Ballpark Check

One liter of 12.0 M HCl solution contains 12 mol of HCl, so 300 mL (0.3 L) of solution contains $0.3 \times 12 = 3.6$ mol.

WORKED EXAMPLE 6.12 Calculating the Number of Moles of Solute in a Solution

Continued

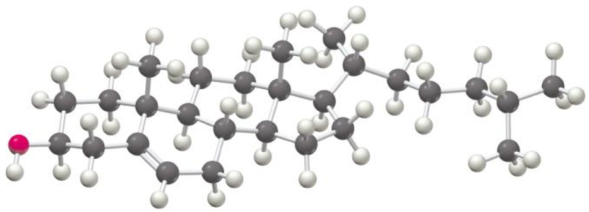
PROBLEM 6.12 How many moles of solute are present in the following solutions?
(a) 125 mL of 0.20 M $NaHCO_3$ **(b)** 650.0 mL of 2.50 M H_2SO_4

PROBLEM 6.13 How many grams of solute would you use to prepare the following solutions?
(a) 500.0 mL of 1.25 M NaOH **(b)** 1.50 L of 0.250 M glucose ($C_6H_{12}O_6$)

PROBLEM 6.14 How many milliliters of a 0.20 M glucose ($C_6H_{12}O_6$) solution are needed to provide a total of 25.0 g of glucose?

PROBLEM 6.15 The concentration of cholesterol ($C_{27}H_{46}O$) in normal blood is approximately 0.005 M. How many grams of cholesterol are in 750 mL of blood?

Cholesterol



WORKED EXAMPLE 6.13 Diluting a Solution

How would you prepare 500.0 mL of 0.2500 M NaOH solution starting from a concentration of 1.000 M?

Strategy

The problem gives initial and final concentrations (M_i and M_f) and final volume (V_f) and asks for the initial volume (V_i) that we need to dilute. Rewriting the equation $M_i \times V_i = M_f \times V_f$ as $V_i = (M_f/M_i) \times V_f$ gives the answer.

Solution

$$V_i = \frac{M_f}{M_i} \times V_f = \frac{0.2500\ \cancel{M}}{1.000\ \cancel{M}} \times 500.0\ \text{mL} = 125.0\ \text{mL}$$

We need to place 125.0 mL of 1.000 M NaOH solution in a 500.0 mL volumetric flask and fill to the mark with water.

Ballpark Check

Because the concentration decreases by a factor of four after dilution (from 1.000 M to 0.2500 M), the volume must increase by a factor of four. Thus, to prepare 500.0 mL of solution, we should start with 500.0/4 = 125.0 mL.

PROBLEM 6.16 What is the final concentration if 75.0 mL of a 3.50 M glucose solution is diluted to a volume of 400.0 mL?

PROBLEM 6.17 Sulfuric acid is normally purchased at a concentration of 18.0 M. How would you prepare 250.0 mL of 0.500 M aqueous H_2SO_4? (Remember to add the acid to water rather than water to the acid.)

WORKED EXAMPLE 6.14 Reaction Stoichiometry in Solution

Stomach acid, a dilute solution of HCl in water, can be neutralized by reaction with sodium hydrogen carbonate, $NaHCO_3$, according to the equation

$$HCl(aq) + NaHCO_3(aq) \longrightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

How many milliliters of 0.125 M $NaHCO_3$ solution are needed to neutralize 18.0 mL of 0.100 M HCl?

Strategy

Solving stoichiometry problems always requires finding the number of moles of one reactant, using the coefficients of the balanced equation to find the number of moles of the other reactant, and then finding the amount of the other reactant. The flow diagram summarizing the situation was shown in Figure 6.4.

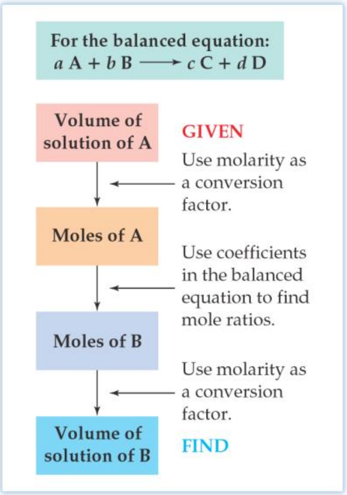


FIGURE 6.4
A flow diagram summarizing the use of molarity as a conversion factor between moles and volume in stoichiometry calculations.

Solution

We first have to find how many moles of HCl are in 18.0 mL of a 0.100 M solution by multiplying volume times molarity:

$$\text{Moles of HCl} = 18.0\ \cancel{\text{mL}} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \frac{0.100\ \text{mol}}{1\ \cancel{\text{L}}} = 1.80 \times 10^{-3}\ \text{mol HCl}$$

Next, check the coefficients of the balanced equation to find that 1 mol of HCl reacts with 1 mol of $NaHCO_3$, and then calculate how many milliliters of a 0.125 M $NaHCO_3$ solution contains 1.80×10^{-3} mol:

$$1.80 \times 10^{-3}\ \cancel{\text{mol HCl}} \times \frac{1\ \cancel{\text{mol NaHCO}_3}}{1\ \cancel{\text{mol HCl}}} \times \frac{1\ \text{L solution}}{0.125\ \cancel{\text{mol NaHCO}_3}} = 0.0144\ \text{L solution}$$

Thus, 14.4 mL of the 0.125 M $NaHCO_3$ solution is needed to neutralize 18.0 mL of the 0.100 M HCl solution.

Ballpark Check

The balanced equation shows that HCl and $NaHCO_3$ react in a 1 : 1 molar ratio, and we are told that the concentrations of the two solutions are about the same. Thus, the volume of the $NaHCO_3$ solution must be about the same as that of the HCl solution.

WORKED EXAMPLE 6.14 Reaction Stoichiometry in Solution

Continued

PROBLEM 6.18 What volume of 0.250 M H_2SO_4 is needed to react with 50.0 mL of 0.100 M NaOH? The equation is

$$H_2SO_4(aq) + 2\,NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2\,H_2O(l)$$

PROBLEM 6.19 What is the molarity of an HNO_3 solution if 68.5 mL is needed to react with 25.0 mL of 0.150 M KOH solution? The equation is

$$HNO_3(aq) + KOH(aq) \longrightarrow KNO_3(aq) + H_2O(l)$$

WORKED EXAMPLE 6.15 Calculating an Empirical Formula from a Percent Composition

Vitamin C (ascorbic acid) contains 40.92% C, 4.58% H, and 54.50% O by mass. What is the empirical formula of ascorbic acid?

Strategy

Assume that you have 100.00 g of ascorbic acid, and then carry out the procedure outlined in Figure 6.7.

Solution

First, find the number of moles of each element in the sample:

$$40.92\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.0\ \cancel{\text{g C}}} = 3.41\ \text{mol C}$$

$$4.58\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.01\ \cancel{\text{g H}}} = 4.53\ \text{mol H}$$

$$54.50\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.0\ \cancel{\text{g O}}} = 3.41\ \text{mol O}$$

Dividing each of the three numbers by the smallest (3.41 mol) gives a C : H : O mole ratio of 1 : 1.33 : 1 and a temporary formula of $C_1H_{1.33}O_1$. Multiplying the subscripts by small integers in a trial-and-error procedure until whole numbers are found then gives the empirical formula: $C_{(3\times1)}H_{(3\times1.33)}O_{(3\times1)} = C_3H_4O_3$.

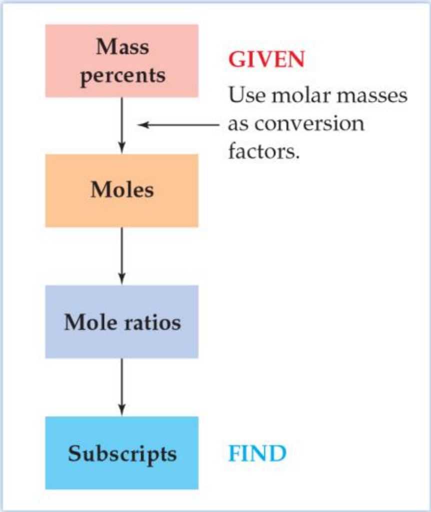


FIGURE 6.7
A flow diagram for calculating the formula of a compound from its percent composition.

WORKED EXAMPLE 6.16 Calculating a Percent Composition from a Formula

Glucose, or blood sugar, has the molecular formula $C_6H_{12}O_6$. What is the empirical formula, and what is the percent composition of glucose?

Strategy and Solution

The percent composition of glucose can be calculated either from the molecular formula ($C_6H_{12}O_6$) or from the empirical formula (CH_2O). Using the molecular formula, for instance, the C : H : O mole ratio of 6 : 12 : 6 can be converted into a mass ratio by assuming that we have 1 mol of compound and carrying out mole-to-gram conversions:

$$1\ \cancel{\text{mol glucose}} \times \frac{6\ \cancel{\text{mol C}}}{1\ \cancel{\text{mol glucose}}} \times \frac{12.0\ \text{g C}}{1\ \cancel{\text{mol C}}} = 72.0\ \text{g C}$$

$$1\ \cancel{\text{mol glucose}} \times \frac{12\ \cancel{\text{mol H}}}{1\ \cancel{\text{mol glucose}}} \times \frac{1.01\ \text{g H}}{1\ \cancel{\text{mol H}}} = 12.1\ \text{g H}$$

$$1\ \cancel{\text{mol glucose}} \times \frac{6\ \cancel{\text{mol O}}}{1\ \cancel{\text{mol glucose}}} \times \frac{16.0\ \text{g O}}{1\ \cancel{\text{mol O}}} = 96.0\ \text{g O}$$

Dividing the mass of each element by the total mass and multiplying by 100% gives the percent composition. Note that the sum of the mass percentages is 100%.

$$\text{Total mass of 1 mol glucose} = 72.0\ \text{g} + 12.1\ \text{g} + 96.0\ \text{g} = 180.1\ \text{g}$$

$$\%\ \text{C} = \frac{72.0\ \text{g C}}{180.1\ \text{g}} \times 100\% = 40.0\%$$

$$\%\ \text{H} = \frac{12.1\ \text{g H}}{180.1\ \text{g}} \times 100\% = 6.72\%$$

$$\%\ \text{O} = \frac{96.0\ \text{g O}}{180.1\ \text{g}} \times 100\% = 53.3\%$$

WORKED EXAMPLE 6.16 Calculating a Percent Composition from a Formula

Continued

PROBLEM 6.22 What is the empirical formula, and what is the percent composition of dimethylhydrazine, $C_2H_8N_2$, a colorless liquid used as a rocket fuel?

PROBLEM 6.23 What is the empirical formula of an ingredient in Bufferin tablets that has the percent composition C 14.25%, O 56.93%, Mg 28.83% by mass?

PROBLEM 6.24 What is the percent composition of citric acid, an organic acid commonly found in citrus fruits? (Gray = C, red = O, H = Ivory.)

WORKED EXAMPLE 6.17 Calculating an Empirical Formula and a Molecular Formula from a Combustion Analysis

Caproic acid, the substance responsible for the aroma of goats, dirty socks, and running shoes, contains carbon, hydrogen, and oxygen. On combustion analysis, a 0.450 g sample of caproic acid gives 0.418 g of H_2O and 1.023 g of CO_2. What is the empirical formula of caproic acid? If the molecular mass of caproic acid is 116.2 amu, what is the molecular formula?

Strategy

Using the steps outlined in Figure 6.8, find the empirical formula of caproic acid, calculate a formula mass, and compare it with the known molecular mass.

Solution

First, find the molar amounts of C and H in the sample:

$$\text{Moles of C} = 1.023\ \cancel{\text{g CO}_2} \times \frac{1\ \cancel{\text{mol CO}_2}}{44.01\ \cancel{\text{g CO}_2}} \times \frac{1\ \text{mol C}}{1\ \cancel{\text{mol CO}_2}} = 0.023\,24\ \text{mol C}$$

$$\text{Moles of H} = 0.418\ \cancel{\text{g H}_2\text{O}} \times \frac{1\ \cancel{\text{mol H}_2\text{O}}}{18.02\ \cancel{\text{g H}_2\text{O}}} \times \frac{2\ \text{mol H}}{1\ \cancel{\text{mol H}_2\text{O}}} = 0.0464\ \text{mol H}$$

Next, find the number of grams of C and H in the sample:

$$\text{Mass of C} = 0.023\,24\ \cancel{\text{mol C}} \times \frac{12.01\ \text{g C}}{1\ \cancel{\text{mol C}}} = 0.2791\ \text{g C}$$

$$\text{Mass of H} = 0.0464\ \cancel{\text{mol H}} \times \frac{1.01\ \text{g H}}{1\ \cancel{\text{mol H}}} = 0.0469\ \text{g H}$$

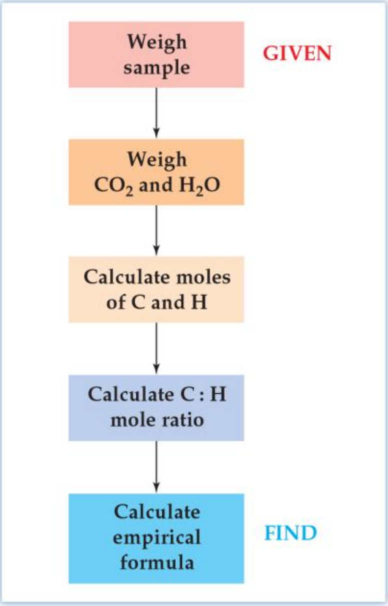


FIGURE 6.8
A flow diagram for determining an empirical formula from combustion analysis of a compound containing C and H.

WORKED EXAMPLE 6.17 Calculating an Empirical Formula and a Molecular Formula from a Combustion Analysis

Continued

Subtracting the masses of C and H from the mass of the starting sample indicates that 0.124 g is unaccounted for:

$$0.450\ \text{g} - (0.2791\ \text{g} + 0.0469\ \text{g}) = 0.124\ \text{g}$$

Because we are told that oxygen is also present in the sample, the "missing" mass must be due to oxygen, which can't be detected by combustion. We therefore need to find the number of moles of oxygen in the sample:

$$\text{Moles of O} = 0.124\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 0.007\ 75\ \text{mol O}$$

Knowing the relative numbers of moles of all three elements, C, H, and O, we divide the three numbers of moles by the smallest number (0.007 75 mol of oxygen) to arrive at a C : H : O ratio of 3 : 6 : 1.

$$C_{\left(\frac{0.02324}{0.00775}\right)}H_{\left(\frac{0.0464}{0.00775}\right)}O_{\left(\frac{0.00775}{0.00775}\right)} = C_3H_6O$$

The empirical formula of caproic acid is therefore C_3H_6O, and the empirical formula mass is 58.1 amu. Because the molecular mass of caproic acid is 116.2, or twice the empirical formula mass, the molecular formula of caproic acid must be $C_{(2\times3)}H_{(2\times6)}O_{(2\times1)} = C_6H_{12}O_2$.

WORKED EXAMPLE 6.17 Calculating an Empirical Formula and a Molecular Formula from a Combustion Analysis

Continued

PROBLEM 6.25 Menthol, a flavoring agent obtained from peppermint oil, contains carbon, hydrogen, and oxygen. On combustion analysis, 1.00 g of menthol yields 1.161 g of H_2O and 2.818 g of CO_2. What is the empirical formula of menthol?

PROBLEM 6.26 Ribose, a sugar present in the cells of all living organisms, has a molecular mass of 150 amu and the empirical formula CH_2O. What is the molecular formula of ribose?

PROBLEM 6.27 Convert the following percent compositions into molecular formulas:

(a) Diborane: H 21.86%, B 78.14%; Molec. Mass = 27.7 amu
(b) Trioxan: C 40.00%, H 6.71%, O 53.28%; Molec. Mass = 90.08 amu