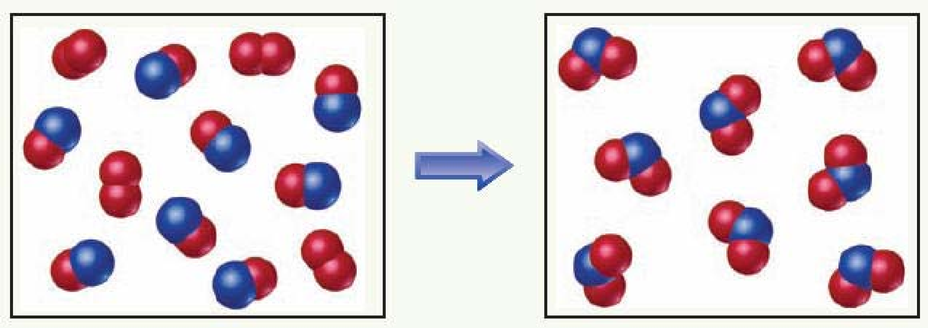
Sample Exercise 3.1 Interpreting and Balancing Chemical Equations

The following diagram represents a chemical reaction in which the red spheres are oxygen atoms and the blue spheres are nitrogen atoms. (a) Write the chemical formulas for the reactants and products. (b) Write a balanced equation for the reaction. (c) Is the diagram consistent with the law of conservation of mass?



Solution

(a) The left box, which represents the reactants, contains two kinds of molecules, those composed of two oxygen atoms (O_2) and those composed of one nitrogen atom and one oxygen atom (NO). The right box, which represents the products, contains only molecules composed of one nitrogen atom and two oxygen atoms (NO_2).

(b) The unbalanced chemical equation is

$$O_2 + NO \rightarrow NO_2 \text{ (unbalanced)}$$

This equation has three O atoms on the left side of the arrow and two O atoms on the right side. We can increase the number of O atoms by placing a coefficient 2 on the product side:

$$O_2 + NO \rightarrow 2\ NO_2 \text{ (unbalanced)}$$

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Sample Exercise 3.1 Interpreting and Balancing Chemical Equations

Solution (continued)

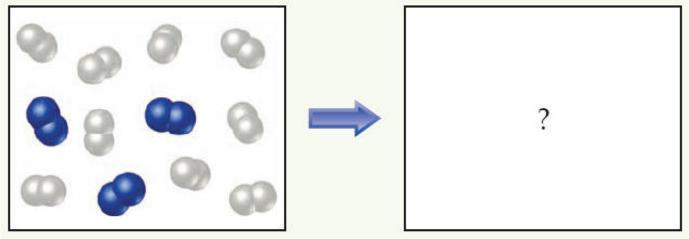
Now there are two N atoms and four O atoms on the right. Placing the coefficient 2 in front of NO balances both the number of N atoms and O atoms:

$$O_2 + 2\ NO \rightarrow 2\ NO_2 \text{ (balanced)}$$

(c) The left box (reactants) contains four O_2 molecules and eight NO molecules. Thus, the molecular ratio is one O_2 for each two NO as required by the balanced equation. The right box (products) contains eight NO_2 molecules. The number of NO_2 molecules on the right equals the number of NO molecules on the left as the balanced equation requires. Counting the atoms, we find eight N atoms in the eight NO molecules in the box on the left. There are also $4 \times 2 = 8$ O atoms in the O_2 molecules and eight O atoms in the NO molecules, giving a total of 16 O atoms. In the box on the right, we find eight N atoms and $8 \times 2 = 16$ O atoms in the eight NO_2 molecules. Because there are equal numbers of both N and O atoms in the two boxes, the drawing is consistent with the law of conservation of mass.

Practice Exercise

In the following diagram, the white spheres represent hydrogen atoms, and the blue spheres represent nitrogen atoms. To be consistent with the law of conservation of mass, how many NH3 molecules should be shown in the right box?



Answer: Six NH_3 molecules.

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Sample Exercise 3.2 Balancing Chemical Equations

Balance this equation:

$$Na(s) + H_2O(l) \rightarrow NaOH(aq) + H_2(g)$$

Solution

Begin by counting each kind of atom on both sides of the arrow. The Na and O atoms are balanced—one Na and one O on each side—but there are two H atoms on the left and three H atoms on the right. Thus, we need to increase the number of H atoms on the left. To begin balancing H, let's try placing the coefficient 2 in front of H_2O:

$$Na(s) + 2\ H_2O(l) \rightarrow 2\ NaOH(aq) + H_2(g)$$

Beginning this way does not balance H but does increase the number of H atoms among the reactants, which we need to do. Adding the coefficient 2 causes O to be unbalanced; we will take care of that after we balance H. Now that we have 2 H_2O on the left, we can balance H by putting the coefficient 2 in front of NaOH on the right:

$$Na(s) + 2\ H_2O(l) \rightarrow 2\ NaOH(aq) + H_2(g)$$

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Sample Exercise 3.2 Balancing Chemical Equations

Solution (continued)

Balancing H in this way fortuitously brings O into balance. But notice that Na is now unbalanced, with one Na on the left and two on the right. To rebalance Na, we put the coefficient 2 in front of the reactant:

$$2Na(s) + 2\ H_2O(l) \rightarrow 2\ NaOH(aq) + H_2(g)$$

Finally, we check the number of atoms of each element and find that we have two Na atoms, four H atoms, and two O atoms on each side of the equation. The equation is balanced.

Comment Notice that in balancing this equation, we moved back and forth placing a coefficient in front of H_2O, then NaOH, and finally Na. In balancing equations, we often find ourselves following this pattern of moving back and forth from one side of the arrow to the other, placing coefficients first in front of a formula on one side and then in front of a formula on the other side until the equation is balanced. You can always tell if you have balanced your equation correctly, no matter how you did it, by checking that the number of atoms of each element is the same on both sides of the arrow.

Practice Exercise

Balance the following equations by providing the missing coefficients:

(a) $\_Fe(s) + \_O_2(g) \rightarrow \_Fe_2O_3(s)$

(b) $\_C_2H_4(g) + \_O_2(g) \rightarrow \_CO_2(g) + \_H_2O(g)$

(c) $\_Al(s) + \_HCl(aq) \rightarrow \_AlCl_3(aq) + \_H2(g)$

Answers: **(a)** 4, 3, 2; **(b)** 1, 3, 2, 2; **(c)** 2, 6, 2, 3

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Sample Exercise 3.3 Writing Balanced Equations for Combination and Decomposition Reactions

Write balanced equations for the following reactions: **(a)** The combination reaction that occurs when lithium metal and fluorine gas react. **(b)** The decomposition reaction that occurs when solid barium carbonate is heated. (Two products form: a solid and a gas.)

Solution

(a) The symbol for lithium is Li. With the exception of mercury, all metals are solids at room temperature. Fluorine occurs as a diatomic molecule (see Figure 2.19). Thus, the reactants are Li(*s*) and $F_2(g)$. The product will be composed of a metal and a nonmetal, so we expect it to be an ionic solid. Lithium ions have a 1+ charge, Li^+, whereas fluoride ions have a 1– charge, F^-.Thus, the chemical formula for the product is LiF. The balanced chemical equation is

$$2\ Li(s) + F_2(g) \rightarrow 2\ LiF(s)$$

(b) The chemical formula for barium carbonate is $BaCO_3$. As noted in the text, many metal carbonates decompose to form metal oxides and carbon dioxide when heated. In Equation 3.7, for example, $CaCO_3$ decomposes to form CaO and CO_2. Thus, we would expect that $BaCO_3$ decomposes to form BaO and CO_2. Barium and calcium are both in group 2A in the periodic table, which further suggests they would react in the same way:

$$BaCO_3(s) \rightarrow BaO(s) + CO_2(g)$$

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Sample Exercise 3.3 Writing Balanced Equations for Combination and Decomposition Reactions

Practice Exercise

Write balanced chemical equations for the following reactions: **(a)** Solid mercury(II) sulfide decomposes into its component elements when heated. **(b)** The surface of aluminum metal undergoes a combination reaction with oxygen in the air.

Answer: (a) $HgS(s) \rightarrow Hg(l) + S(s)$ **(b)** $4\ Al(s) + 3\ O_2(g) \rightarrow 2\ Al_2O_3(s)$

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Sample Exercise 3.4 Writing Equations for Combustion Reactions

Write the balanced equation for the reaction that occurs when methanol, $CH_3OH(l)$, is burned in air.

Solution

When any compound containing C, H, and O is combusted, it reacts with the $O_2(g)$ in air to produce $CO_2(g)$ and $H_2O(g)$. Thus, the unbalanced equation is

$$CH_3OH(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

In this equation the C atoms are balanced with one carbon on each side of the arrow. Because CH_3OH has four H atoms, we place the coefficient 2 in front of H_2O to balance the H atoms:

$$CH_3OH(l) + O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$$

Adding the coefficient balances H but gives four O atoms in the products. Because there are only three O atoms in the reactants (one in CH_3OH and two in O_2), we are not finished yet. We can place the fractional coefficient 2/3 in front of O_2 to give a total of four O atoms in the reactants (there are 2/3 × 2 = 3 O atoms in 3/2 O_2): $CH_3OH(l) + \frac{3}{2}\ O_2(g) \longrightarrow CO_2(g) + 2\ H_2O(g)$

Although the equation is now balanced, it is not in its most conventional form because it contains a fractional coefficient. If we multiply each side of the equation by 2, we will remove the fraction and achieve the following balanced equation:

$$2\ CH_3OH(l) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 4\ H_2O(g)$$

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Practice Exercise

Write the balanced equation for the reaction that occurs when ethanol, $C_2H_5OH(l)$, is burned in air.

Answer: $C_2H_5OH(l) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(g)$

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Sample Exercise 3.5 Calculating formula Weights

Calculate the formula weight of **(a)** sucrose, $C1_2H_{22}O_{11}$ (table sugar), and **(b)** calcium nitrate, $Ca(NO_3)_2$.

Solution

(a) By adding the atomic weights of the atoms in sucrose, we find the formula weight to be 342.0 amu:

$$\begin{aligned} 12\text{ C atoms} &= 12(12.0\text{ amu}) = 144.0\text{ amu} \\ 22\text{ H atoms} &= 22(1.0\text{ amu}) = 22.0\text{ amu} \\ 11\text{ O atoms} &= 11(16.0\text{ amu}) = \underline{176.0\text{ amu}} \\ & \qquad\qquad\qquad\quad 342.0\text{ amu} \end{aligned}$$

(b) If a chemical formula has parentheses, the subscript outside the parentheses is a multiplier for all atoms inside. Thus, for Ca(NO3)2, we have

$$\begin{aligned} 1\text{ Ca atom} &= 1(40.1\text{ amu}) = 40.1\text{ amu} \\ 2\text{ N atoms} &= 2(14.0\text{ amu}) = 28.0\text{ amu} \\ 6\text{ O atoms} &= 6(16.0\text{ amu}) = \underline{96.0\text{ amu}} \\ & \qquad\qquad\qquad\quad 164.1\text{ amu} \end{aligned}$$

Practice Exercise

Calculate the formula weight of **(a)** $Al(OH)_3$ and **(b)** CH_3OH.
Answer: **(a)** 78.0 amu, **(b)** 32.0 amu

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Sample Exercise 3.6 Calculating Percentage Composition

Calculate the percentage of carbon, hydrogen, and oxygen (by mass) in $C_{12}H_{22}O_{11}$.

Solution

Let's examine this question using the problem-solving steps in the "Strategies in Chemistry: Problem Solving" essay that appears on the next page.

Analyze We are given a chemical formula, $C_{12}H_{22}O_{11}$, and asked to calculate the percentage by mass of its component elements (C, H, and O).

Plan We can use Equation 3.10, relying on a periodic table to obtain the atomic weight of each component element. The atomic weights are first used to determine the formula weight of the compound. (The formula weight of $C_{12}H_{22}O_{11}$, 342.0 amu, was calculated in Sample Exercise 3.5.) We must then do three calculations, one for each element.

Solve Using Equation 3.10, we have

$$\%C = \frac{(12)(12.0\text{ amu})}{342.0\text{ amu}} \times 100\% = 42.1\%$$

$$\%H = \frac{(22)(1.0\text{ amu})}{342.0\text{ amu}} \times 100\% = 6.4\%$$

$$\%O = \frac{(11)(16.0\text{ amu})}{342.0\text{ amu}} \times 100\% = 51.5\%$$

Check The percentages of the individual elements must add up to 100%, which they do in this case. We could have used more significant figures for our atomic weights, giving more significant figures for our percentage composition, but we have adhered to our suggested guideline of rounding atomic weights to one digit beyond the decimal point.

Practice Exercise

Calculate the percentage of nitrogen, by mass, in $Ca(NO_3)_2$.

Answer: 17.1%

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Sample Exercise 3.7 Estimating Numbers in Atoms

Without using a calculator, arrange the following samples in order of increasing numbers of carbon atoms: 12 g ^{12}C, 1 mol C_2H_2, 9×10^{23} molecules of CO_2.

Solution

Analyze We are given amounts of different substances expressed in grams, moles, and number of molecules and asked to arrange the samples in order of increasing numbers of C atoms.

Plan To determine the number of C atoms in each sample, we must convert g ^{12}C, 1 mol C_2H_2, and 9×10^{23} molecules CO_2 all to numbers of C atoms. To make these conversions, we use the definition of mole and Avogadro's number.

Solve A mole is defined as the amount of matter that contains as many units of the matter as there are C atoms in exactly 12 g of ^{12}C. Thus, 12 g of ^{12}C contains 1 mol of C atoms (that is, 6.02×10^{23} C atoms). One mol of C_2H_2 contains 6×10^{23} C_2H_2 molecules. Because there are two C atoms in each C_2H_2 molecule, this sample contains 12×10^{23} C atoms. Because each CO_2 molecule contains one C atom, the sample of CO_2 contains 9×10^{23} C atoms. Hence, the order is 12 g ^{12}C (6×10^{23} C atoms) < 9×10^{23} CO_2 molecules (9×10^{23} C atoms) < 1 mol C_2H_2 (12×10^{23} C atoms).

Check We can check our results by comparing the number of moles of C atoms in each sample because the number of moles is proportional to the number of atoms. Thus, 12 g of ^{12}C is 1 mol C; 1 mol of C_2H_2 contains 2 mol C, and 9×10^{23} molecules of CO_2 contain 1.5 mol C, giving the same order as above: 12 g ^{12}C (1 mol C) < 9×10^{23} CO_2 molecules (1.5 mol C) < 1 mol C_2H_2 (2 mol C).

Practice Exercise

Without using a calculator, arrange the following samples in order of increasing number of O atoms: 1 mol H_2O, 1 mol CO_2, 3×10^{23} molecules O_3.

Answer: 1 mol H_2O (6×10^{23} O atoms) 3×10^{23} molecules O_3 (9×10^{23} O atoms) 1 mol CO_2 (12×10^{23} O atoms)

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Sample Exercise 3.8 Converting Moles to Atoms

Calculate the number of H atoms in 0.350 mol of $C_6H_{12}O_6$.

Solution

Analyze We are given both the amount of a substance (0.350 mol) and its chemical formula ($C_6H_{12}O_6$). The unknown is the number of H atoms in the sample.

Plan Avogadro's number provides the conversion factor between the number of moles of $C_6H_{12}O_6$ and the number of molecules of $C_6H_{12}O_6$. Once we know the number of molecules of $C_6H_{12}O_6$, we can use the chemical formula, which tells us that each molecule of $C_6H_{12}O_6$ contains 12 H atoms. Thus, we convert moles of $C_6H_{12}O_6$ to molecules of $C_6H_{12}O_6$ and then determine the number of atoms of H from the number of molecules of $C_6H_{12}O_6$:

$$\text{Moles } C_6H_{12}O_6 \rightarrow \text{molecules } C_6H_{12}O_6 \rightarrow \text{atoms H}$$

Solve

$$\begin{aligned}\text{H atoms} &= (0.350\ \cancel{\text{mol}\ C_6H_{12}O_6})\left(\frac{6.02\times10^{23}\ \cancel{\text{molecules}\ C_6H_{12}O_6}}{1\ \cancel{\text{mol}\ C_6H_{12}O_6}}\right)\left(\frac{12\ \text{H atoms}}{1\ \cancel{\text{molecule}\ C_6H_{12}O_6}}\right)\\ &= 2.53\times10^{24}\ \text{H atoms}\end{aligned}$$

Check The magnitude of our answer is reasonable. It is a large number about the magnitude of Avogadro's number. We can also make the following ballpark calculation: Multiplying $0.35 \times 6 \times 10^{23}$ gives about 2×10^{23} molecules. Multiplying this result by 12 gives $24 \times 10^{23} = 2.4 \times 10^{24}$ H atoms, which agrees with the previous, more detailed calculation. Because we were asked for the number of H atoms, the units of our answer are correct. The given data had three significant figures, so our answer has three significant figures.

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Sample Exercise 3.8 Converting Moles to Atoms

Calculate the number of H atoms in 0.350 mol of $C_6H_{12}O_6$.

Practice Exercise

How many oxygen atoms are in **(a)** 0.25 mol $Ca(NO_3)_2$ and **(b)** 1.50 mol of sodium carbonate?
Answer: **(a)** 9.0×10^{23}, **(b)** 2.71×10^{24}

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Sample Exercise 3.9 Calculating Molar Mass

What is the mass in grams of 1.000 mol of glucose, $C_6H_{12}O_6$?

Solution

Analyze We are given a chemical formula and asked to determine its molar mass.
Plan The molar mass of a substance is found by adding the atomic weights of its component atoms.
Solve

$$\begin{aligned} 6\text{ C atoms} &= 6(12.0\text{ amu}) = 72.0\text{ amu} \\ 12\text{ H atoms} &= 12(1.0\text{ amu}) = 12.0\text{ amu} \\ 6\text{ O atoms} &= 6(16.0\text{ amu}) = \underline{96.0\text{ amu}} \\ & \qquad\qquad\qquad\qquad 180.0\text{ amu} \end{aligned}$$

Because glucose has a formula weight of 180.0 amu, one mole of this substance has a mass of 180.0 g. In other words, $C_6H_{12}O_6$ has a molar mass of 180.0 g/mol.
Check The magnitude of our answer seems reasonable, and g/mol is the appropriate unit for the molar mass.
Comment Glucose is sometimes called dextrose. Also known as blood sugar, glucose is found widely in nature, occurring in honey and fruits. Other types of sugars used as food are converted into glucose in the stomach or liver before the body uses them as energy sources. Because glucose requires no conversion, it is often given intravenously to patients who need immediate nourishment. People who have diabetes must carefully monitor the amount of glucose in their blood (See “Chemistry and Life” box in Section 3.6).

Practice Exercise

Calculate the molar mass of $Ca(NO_3)_2$.
Answer: 164.1 g/mol

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Sample Exercise 3.10 Converting Grams to Moles

Calculate the number of moles of glucose ($C_6H_{12}O_6$) in 5.380 g of $C_6H_{12}O_6$.

Solution

Analyze We are given the number of grams of a substance and its chemical formula and asked to calculate the number of moles.

Plan The molar mass of a substance provides the factor for converting grams to moles. The molar mass of $C_6H_{12}O_6$ is 180.0 g/mol (Sample Exercise 3.9).

Solve Using 1 mol $C_6H_{12}O_6$=180.0 g $C_6H_{12}O_6$ to write the appropriate conversion factor, we have

$$\text{Moles } C_6H_{12}O_6 = (5.380\ \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1 \text{ mol } C_6H_{12}O_6}{180.0\ \cancel{\text{g } C_6H_{12}O_6}}\right) = 0.02989 \text{ mol } C_6H_{12}O_6$$

Check Because 5.380 g is less than the molar mass, a reasonable answer is less than one mole. The units of our answer (mol) are appropriate. The original data had four significant figures, so our answer has four significant figures.

Practice Exercise

How many moles of sodium bicarbonate ($NaHCO_3$) are in 508 g of $NaHCO_3$?

Answer: 6.05 mol $NaHCO_3$

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Sample Exercise 3.11 Converting Moles to Grams

Calculate the mass, in grams, of 0.433 mol of calcium nitrate.

Solution

Analyze We are given the number of moles and the name of a substance and asked to calculate the number of grams in the sample.
Plan To convert moles to grams, we need the molar mass, which we can calculate using the chemical formula and atomic weights.
Solve Because the calcium ion is Ca^{2+} and the nitrate ion is NO_3^-, calcium nitrate is $Ca(NO_3)_2$. Adding the atomic weights of the elements in the compound gives a formula weight of 164.1 amu. Using 1 mol $Ca(NO_3)_2$ = 164.1 g $Ca(NO_3)_2$ to write the appropriate conversion factor, we have

$$\text{Grams } Ca(NO_3)_2 = (0.433\ \cancel{\text{mol } Ca(NO_3)_2})\left(\frac{164.1\text{ g } Ca(NO_3)_2}{1\ \cancel{\text{mol } Ca(NO_3)_2}}\right) = 71.1\text{ g } Ca(NO_3)_2$$

Check The number of moles is less than 1, so the number of grams must be less than the molar mass, 164.1 g. Using rounded numbers to estimate, we have 0.5 × 150 = 75 g. The magnitude of our answer is reasonable. Both the units (g) and the number of significant figures (3) are correct.

Practice Exercise

What is the mass, in grams, of **(a)** 6.33 mol of $NaHCO_3$ and **(b)** 3.0×10^{-5} mol of sulfuric acid?
Answer: **(a)** 532 g, **(b)** 2.9×10^{-3} g

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Sample Exercise 3.12 Calculating the Number of Molecules and Number of Atoms from Mass

(a) How many glucose molecules are in 5.23 g of $C_6H_{12}O_6$? **(b)** How many oxygen atoms are in this sample?

Solution

Analyze We are given the number of grams and the chemical formula and asked to calculate (a) the number of molecules and (b) the number of O atoms in the sample.

(a) Plan The strategy for determining the number of molecules in a given quantity of a substance is summarized in Figure 3.10. We must convert 5.23 g $C_6H_{12}O_6$ to moles $C_6H_{12}O_6$, which can then be converted to molecules $C_6H_{12}O_6$. The first conversion uses the molar mass of $C_6H_{12}O_6$: 1 mol $C_6H_{12}O_6$ = 180.0 g $C_6H_{12}O_6$. The second conversion uses Avogadro's number.

Solve

Molecules $C_6H_{12}O_6$

$$= (5.23\ \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1\ \cancel{\text{mol } C_6H_{12}O_6}}{180.0\ \cancel{\text{g } C_6H_{12}O_6}}\right)\left(\frac{6.02 \times 10^{23} \text{ molecules } C_6H_{12}O_6}{1\ \cancel{\text{mol } C_6H_{12}O_6}}\right)$$

$$= 1.75 \times 10^{22} \text{ molecules } C_6H_{12}O_6$$

Check The magnitude of the answer is reasonable. Because the mass we began with is less than a mole, there should be fewer than 6.02×10^{23} molecules. We can make a ballpark estimate of the answer: 5/200 = 2.5 10^{-2} mol; $2.5 \times 10^{-2} \times 6 \times 10^{23} = 15 \times 10^{21} = 1.5 \times 10^{22}$ molecules. The units (molecules) and significant figures (three) are appropriate.

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Sample Exercise 3.12 Calculating the Number of Molecules and Number of Atoms from Mass

Solution (continued)

(b) Plan To determine the number of O atoms, we use the fact that there are six O atoms in each molecule of $C_6H_{12}O_6$. Thus, multiplying the number of molecules $C_6H_{12}O_6$ by the factor (6 atoms O/1 molecule $C_6H_{12}O_6$) gives the number of O atoms.

Solve

$$\text{Atoms O} = (1.75 \times 10^{22}\ \cancel{\text{molecules } C_6H_{12}O_6})\left(\frac{6\ \text{atoms O}}{1\ \cancel{\text{molecule } C_6H_{12}O_6}}\right)$$
$$= 1.05 \times 10^{23}\ \text{atoms O}$$

Check The answer is simply 6 times as large as the answer to part (a). The number of significant figures (three) and the units (atoms O) are correct.

Practice Exercise

(a) How many nitric acid molecules are in 4.20 g of HNO_3? **(b)** How many O atoms are in this sample?
Answer: **(a)** 4.01×10^{22} molecules HNO_3, **(b)** 1.20×10^{23} atoms O

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Sample Exercise 3.13 Calculating Empirical Formula

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

Solution

Analyze We are to determine an empirical formula of a compound from the mass percentages of its elements.
Plan The strategy for determining the empirical formula involves the three steps given in Figure 3.11.

Solve We *first* assume, for simplicity, that we have exactly 100 g of material (although any mass can be used). In 100 g of ascorbic acid, therefore, we have

40.92 g C, 4.58 g H, and 54.50 g O.

Second, we calculate the number of moles of each element:

$$\text{Moles C} = (40.92\ \cancel{\text{g C}})\left(\frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}}\right) = 3.407\ \text{mol C}$$

$$\text{Moles H} = (4.58\ \cancel{\text{g H}})\left(\frac{1\ \text{mol H}}{1.008\ \cancel{\text{g H}}}\right) = 4.54\ \text{mol H}$$

$$\text{Moles O} = (54.50\ \cancel{\text{g O}})\left(\frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}}\right) = 3.406\ \text{mol O}$$

Third, we determine the simplest whole-number ratio of moles by dividing each number of moles by the smallest number of moles, 3.406:

$$\text{C: } \frac{3.407}{3.406} = 1.000 \qquad \text{H: } \frac{4.54}{3.406} = 1.33 \qquad \text{O: } \frac{3.406}{3.406} = 1.000$$

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Sample Exercise 3.13 Calculating Empirical Formula

Solution (continued)

The ratio for H is too far from 1 to attribute the difference to experimental error; in fact, it is quite close to 1 1/3. This suggests that if we multiply the ratio by 3, we will obtain whole numbers:

$$C{:}H{:}O = 3(1{:}1.33{:}1) = 3{:}4{:}3$$

The whole-number mole ratio gives us the subscripts for the empirical formula:

$$C_3H_4O_3$$

Check It is reassuring that the subscripts are moderately sized whole numbers. Otherwise, we have little by which to judge the reasonableness of our answer.

Practice Exercise

A 5.325-g sample of methyl benzoate, a compound used in the manufacture of perfumes, contains 3.758 g of carbon, 0.316 g of hydrogen, and 1.251 g of oxygen. What is the empirical formula of this substance?

Answer: C_4H_4O

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Sample Exercise 3.14 Determining a Molecular Formula

Mesitylene, a hydrocarbon that occurs in small amounts in crude oil, has an empirical formula of C_3H_4. The experimentally determined molecular weight of this substance is 121 amu. What is the molecular formula of mesitylene?

Solution

Analyze We are given an empirical formula and a molecular weight and asked to determine a molecular formula.

Plan The subscripts in the molecular formula of a compound are whole-number multiples of the subscripts in its empirical formula. To find the appropriate multiple, we must compare the molecular weight with the formula weight of the empirical formula.

Solve First, we calculate the formula weight of the empirical formula, C_3H_4:

$$3(12.0 \text{ amu}) + 4(1.0 \text{ amu}) = 40.0 \text{ amu}$$

Next, we divide the molecular weight by the empirical formula weight to obtain the multiple used to multiply the subscripts in C_3H_4:

$$\frac{\text{Molecular weight}}{\text{Empirical formula weight}} = \frac{121}{40.0} = 3.02$$

Only whole-number ratios make physical sense because we must be dealing with whole atoms. The 3.02 in this case could result from a small experimental error in the molecular weight. We therefore multiply each subscript in the empirical formula by 3 to give the molecular formula: C_9H_{12}.

PEARSON Education
Chemistry: The Central Science, Eleventh Edition
By Theodore E. Brown, H. Eugene LeMay, Bruce E. Bursten, and Catherine J. Murphy
With contributions from Patrick Woodward

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Sample Exercise 3.14 Determining a Molecular Formula

Solution

Check We can have confidence in the result because dividing the molecular weight by the formula weight yields nearly a whole number.

Practice Exercise

Ethylene glycol, the substance used in automobile antifreeze, is composed of 38.7% C, 9.7% H, and 51.6% O by mass. Its molar mass is 62.1 g/mol. **(a)** What is the empirical formula of ethylene glycol? **(b)** What is its molecular formula?

Answers: **(a)** CH_3O, **(b)** $C_2H_6O_2$

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Sample Exercise 3.15 Determining Empirical Formula by Combustion Analysis

Isopropyl alcohol, a substance sold as rubbing alcohol, is composed of C, H, and O. Combustion of 0.255 g of isopropyl alcohol produces 0.561 g of CO_2 and 0.306 g of H_2O. Determine the empirical formula of isopropyl alcohol.

Solution

Analyze We are told that isopropyl alcohol contains C, H, and O atoms and given the quantities of CO_2 and H_2O produced when a given quantity of the alcohol is combusted. We must use this information to determine the empirical formula for isopropyl alcohol, a task that requires us to calculate the number of moles of C, H, and O in the sample.

Plan We can use the mole concept to calculate the number of grams of C present in the CO_2 and the number of grams of H present in the H_2O. These amounts are the quantities of C and H present in the isopropyl alcohol before combustion. The number of grams of O in the compound equals the mass of the isopropyl alcohol minus the sum of the C and H masses. Once we have the number of grams of C, H, and O in the sample, we can then proceed as in Sample Exercise 3.13. We can calculate the number of moles of each element, and determine the mole ratio, which gives the subscripts in the empirical formula.

Solve To calculate the number of grams of C, we first use the molar mass of CO_2, 1 mol CO_2 = 44.0 g CO_2, to convert grams of CO_2 to moles of CO_2. Because each CO_2 molecule has only 1 C atom, there is 1 mol of C atoms per mole of CO_2 molecules. This fact allows us to convert the moles of CO_2 to moles of C. Finally, we use the molar mass of C, 1 mol C =1 2.0 g C, to convert moles of C to grams of C. Combining the three conversion factors, we have

$$\text{Grams C} = (0.561\ \cancel{\text{g}\ CO_2})\left(\frac{1\ \cancel{\text{mol}\ CO_2}}{44.0\ \cancel{\text{g}\ CO_2}}\right)\left(\frac{1\ \cancel{\text{mol C}}}{1\ \cancel{\text{mol}\ CO_2}}\right)\left(\frac{12.0\ \text{g C}}{1\ \cancel{\text{mol C}}}\right) = 0.153\ \text{g C}$$

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Sample Exercise 3.15 Determining Empirical Formula by Combustion Analysis

Solution (continued)

The calculation of the number of grams of H from the grams of H_2O is similar, although we must remember that there are 2 mol of H atoms per 1 mol of H_2O molecules:

$$\text{Grams H} = (0.306\ \cancel{\text{g H}_2\text{O}})\left(\frac{1\ \cancel{\text{mol H}_2\text{O}}}{18.0\ \cancel{\text{g H}_2\text{O}}}\right)\left(\frac{2\ \cancel{\text{mol H}}}{1\ \cancel{\text{mol H}_2\text{O}}}\right)\left(\frac{1.01\ \text{g H}}{1\ \cancel{\text{mol H}}}\right) = 0.0343\ \text{g H}$$

The total mass of the sample, 0.255 g, is the sum of the masses of the C, H, and O. Thus, we can calculate the mass of O as follows:

$$\begin{aligned}\text{Mass of O} &= \text{mass of sample} - (\text{mass of C} + \text{mass of H})\\ &= 0.255\ \text{g} - (0.153\ \text{g} + 0.0343\ \text{g}) = 0.068\ \text{g O}\end{aligned}$$

We then calculate the number of moles of C, H, and O in the sample:

$$\text{Moles C} = (0.153\ \cancel{\text{g C}})\left(\frac{1\ \text{mol C}}{12.0\ \cancel{\text{g C}}}\right) = 0.0128\ \text{mol C}$$

$$\text{Moles H} = (0.0343\ \cancel{\text{g H}})\left(\frac{1\ \text{mol H}}{1.01\ \cancel{\text{g H}}}\right) = 0.0340\ \text{mol H}$$

$$\text{Moles O} = (0.068\ \cancel{\text{g O}})\left(\frac{1\ \text{mol O}}{16.0\ \cancel{\text{g O}}}\right) = 0.0043\ \text{mol O}$$

To find the empirical formula, we must compare the relative number of moles of each element in the sample. The relative number of moles of each element is found by dividing each number by the smallest number, 0.0043. The mole ratio of C:H:O so obtained is 2.98:7.91:1.00. The first two numbers are very close to the whole numbers 3 and 8, giving the empirical formula C_3H_8O.

Check The subscripts work out to be moderately sized whole numbers, as expected.

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Sample Exercise 3.15 Determining Empirical Formula by Combustion Analysis

Practice Exercise

(a) Caproic acid, which is responsible for the foul odor of dirty socks, is composed of C, H, and O atoms. Combustion of a 0.225-g sample of this compound produces 0.512 g CO_2 and 0.209 g H_2O. What is the empirical formula of caproic acid? **(b)** Caproic acid has a molar mass of 116 g/mol. What is its molecular formula?

Answers: **(a)** C_3H_6O, **(b)** $C_6H_{12}O_2$

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Sample Exercise 3.16 Calculating Amounts of Reactants and Products

How many grams of water are produced in the oxidation of 1.00 g of glucose, $C_6H_{12}O_6$?

$$C_6H_{12}O_6(s) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(l)$$

Solution

Analyze We are given the mass of a reactant and are asked to determine the mass of a product in the given equation.

Plan The general strategy, as outlined in Figure 3.13, requires three steps. First, the amount of $C_6H_{12}O_6$ must be converted from grams to moles. Second, we can use the balanced equation, which relates the moles of $C_6H_{12}O_6$ to the moles of H_2O: 1 mol $C_6H_{12}O_6 \simeq$ 6 mol H_2O. Third, we must convert the moles of H_2O to grams.

Solve First, use the molar mass of $C_6H_{12}O_6$ to convert from grams $C_6H_{12}O_6$ to moles $C_6H_{12}O_6$:

$$\text{Moles } C_6H_{12}O_6 = (1.00\ \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1 \text{ mol } C_6H_{12}O_6}{180.0\ \cancel{\text{g } C_6H_{12}O_6}}\right)$$

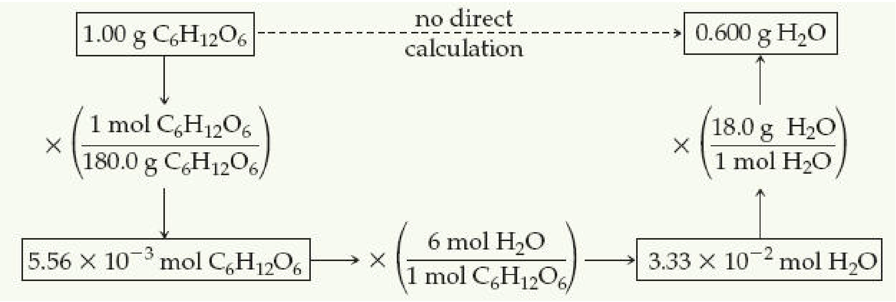
Second, use the balanced equation to convert moles of $C_6H_{12}O_6$ to moles of H_2O:

$$\text{Moles } H_2O = (1.00\ \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1\ \cancel{\text{mol } C_6H_{12}O_6}}{180.0\ \cancel{\text{g } C_6H_{12}O_6}}\right)\left(\frac{6 \text{ mol } H_2O}{1\ \cancel{\text{mol } C_6H_{12}O_6}}\right)$$

Third, use the molar mass of H_2O to convert from moles of H_2O to grams of H_2O:

$$\text{Grams } H_2O = (1.00\ \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1\ \cancel{\text{mol } C_6H_{12}O_6}}{180.0\ \cancel{\text{g } C_6H_{12}O_6}}\right)\left(\frac{6\ \cancel{\text{mol } H_2O}}{1\ \cancel{\text{mol } C_6H_{12}O_6}}\right)\left(\frac{18.0 \text{ g } H_2O}{1\ \cancel{\text{mol } H_2O}}\right)$$
$$= 0.600 \text{ g } H_2O$$

The steps can be summarized in a diagram like that in Figure 3.13:



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Sample Exercise 3.16 Calculating Amounts of Reactants and Products

Solution (continued)

Check An estimate of the magnitude of our answer, 18/180 = 0.1 and 0.1 × 6 = 0.6, agrees with the exact calculation. The units, grams H_2O, are correct. The initial data had three significant figures, so three significant figures for the answer is correct.

Comment An average person ingests 2 L of water daily and eliminates 2.4 L. The difference between 2 L and 2.4 L is produced in the metabolism of foodstuffs, such as in the oxidation of glucose. (*Metabolism* is a general term used to describe all the chemical processes of a living animal or plant.) The desert rat (kangaroo rat), on the other hand, apparently never drinks water. It survives on its metabolic water.

Practice Exercise

The decomposition of $KClO_3$ is commonly used to prepare small amounts of O_2 in the laboratory: $2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$. How many grams of O_2 can be prepared from 4.50 g of $KClO_3$?

Answer: 1.77 g

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Sample Exercise 3.17 Calculating Amounts of Reactants and Products

Solid lithium hydroxide is used in space vehicles to remove the carbon dioxide exhaled by astronauts. The lithium hydroxide reacts with gaseous carbon dioxide to form solid lithium carbonate and liquid water. How many grams of carbon dioxide can be absorbed by 1.00 g of lithium hydroxide?

Solution

Analyze We are given a verbal description of a reaction and asked to calculate the number of grams of one reactant that reacts with 1.00 g of another.
Plan The verbal description of the reaction can be used to write a balanced equation:

$$2\ \mathrm{LiOH}(s) + \mathrm{CO_2}(g) \rightarrow \mathrm{Li_2CO_3}(s) + \mathrm{H_2O}(l)$$

We are given the grams of LiOH and asked to calculate grams of CO_2. We can accomplish this task by using the following sequence of conversions:

$$\text{Grams LiOH} \rightarrow \text{moles LiOH} \rightarrow \text{moles } \mathrm{CO_2} \rightarrow \text{grams CO}$$

The conversion from grams of LiOH to moles of LiOH requires the molar mass of LiOH (6.94 + 16.00 + 1.01 = 23.95 g/mol). The conversion of moles of LiOH to LiOH (6.94 + 16.00 + 1.01 = 23.95 g/mol) moles of CO_2 is based on the balanced chemical equation: 2 mol LiOH $\simeq$ 1 mol CO_2. To convert the number of moles of CO_2 to grams, we must use the molar mass of CO_2: 12.01 + 2(16.00) = 44.01 g/mol.

Solve

$$(1.00\ \cancel{\text{g LiOH}})\left(\frac{1\ \cancel{\text{mol LiOH}}}{23.95\ \cancel{\text{g LiOH}}}\right)\left(\frac{1\ \cancel{\text{mol CO}_2}}{2\ \cancel{\text{mol LiOH}}}\right)\left(\frac{44.01\ \text{g CO}_2}{1\ \cancel{\text{mol CO}_2}}\right) = 0.919\ \text{g CO}_2$$

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Sample Exercise 3.17 Calculating Amounts of Reactants and Products

Solution (continued)

Check Notice that 23.95 ≈ 24, 24 × 2 = 48, and 44/48 is slightly less than 1. The magnitude of the answer is reasonable based on the amount of starting LiOH; the significant figures and units are appropriate, too.

Practice Exercise

Propane, C_3H_8, is a common fuel used for cooking and home heating. What mass of O_2 is consumed in the combustion of 1.00 g of propane?
Answer: 3.64 g

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Sample Exercise 3.18 Calculating the Amount of Product Formed from a Limiting Reactant

The most important commercial process for converting N_2 from the air into nitrogen-containing compounds is based on the reaction of N_2 and H_2 to form ammonia (NH_3):

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

How many moles of NH_3 can be formed from 3.0 mol of N_2 and 6.0 mol of H_2?

Solution

Analyze We are asked to calculate the number of moles of product, NH_3, given the quantities of each reactant, N_2 and H_2, available in a reaction. Thus, this is a limiting reactant problem.

Plan If we assume that one reactant is completely consumed, we can calculate how much of the second reactant is needed in the reaction. By comparing the calculated quantity with the available amount, we can determine which reactant is limiting. We then proceed with the calculation, using the quantity of the limiting reactant.

Solve The number of moles of H_2 needed for complete consumption of 3.0 mol of N_2 is:

$$\text{Moles } H_2 = (3.0\ \cancel{\text{mol } N_2})\left(\frac{3\text{ mol } H_2}{1\ \cancel{\text{mol } N_2}}\right) = 9.0\text{ mol } H_2$$

Because only 6.0 mol H_2 is available, we will run out of H_2 before the N_2 is gone, and H_2 will be the limiting reactant. We use the quantity of the limiting reactant, H_2, to calculate the quantity of NH_3 produced:

$$\text{Moles } NH_3 = (6.0\ \cancel{\text{mol } H_2})\left(\frac{2\text{ mol } NH_3}{3\ \cancel{\text{mol } H_2}}\right) = 4.0\text{ mol } NH_3$$

Comment The table on the right summarizes this example:

	$N_2(g)$	+	$3\,H_2(g)$	$\longrightarrow$	$2\,NH_3(g)$
Initial quantities:	3.0 mol		6.0 mol		0 mol
Change (reaction):	−2.0 mol		−6.0 mol		+4.0 mol
Final quantities:	1.0 mol		0 mol		4.0 mol

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Sample Exercise 3.18 Calculating the Amount of Product Formed from a Limiting Reactant

Solution (continued)

Notice that we can calculate not only the number of moles of NH_3 formed but also the number of moles of each of the reactants remaining after the reaction. Notice also that although the number of moles of H_2 present at the beginning of the reaction is greater than the number of moles of N_2 present, the H_2 is nevertheless the limiting reactant because of its larger coefficient in the balanced equation.

Check The summarizing table shows that the mole ratio of reactants used and product formed conforms to the coefficients in the balanced equation, 1:3:2. Also, because H_2 is the limiting reactant, it is completely consumed in the reaction, leaving 0 mol at the end. Because 6.0 mol H_2 has two significant figures, our answer has two significant figures.

Practice Exercise

Consider the reaction $2\ Al(s) + 3\ Cl_2(g) \rightarrow 2\ AlCl_3(s)$. A mixture of 1.50 mol of Al and 3.00 mol of Cl_2 is allowed to react. **(a)** Which is the limiting reactant? **(b)** How many moles of $AlCl_3$ are formed? **(c)** How many moles of the excess reactant remain at the end of the reaction?

Answers: **(a)** Al, **(b)** 1.50 mol, **(c)** 0.75 mol Cl_2

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Sample Exercise 3.19 Calculating the Amount of Product Formed from a Limiting Reactant

Consider the following reaction that occurs in a fuel cell:

$$2\ H_2(g) + O_2\ (g) \rightarrow 2\ H_2O\ (g)$$

This reaction, properly done, produces energy in the form of electricity and water. Suppose a fuel cell is set up with 150 g of hydrogen gas and 1500 grams of oxygen gas (each measurement is given with two significant figures). How many grams of water can be formed?

Solution

Analyze We are asked to calculate the amount of a product, given the amounts of two reactants, so this is a limiting reactant problem.

Plan We must first identify the limiting reagent. To do so, we can calculate the number of moles of each reactant and compare their ratio with that required by the balanced equation. We then use the quantity of the limiting reagent to calculate the mass of water that forms.

Solve From the balanced equation, we have the following stoichiometric relations:

$$2\ \text{mol}\ H_2 \simeq 1\ \text{mol}\ O_2 \simeq 2\ \text{mol}\ H_2O$$

Using the molar mass of each substance, we can calculate the number of moles of each reactant:

$$\text{Moles}\ H_2 = (150\ \cancel{\text{g}\ H_2})\left(\frac{1\ \text{mol}\ H_2}{2.00\ \cancel{\text{g}\ H_2}}\right) = 75\ \text{mol}\ H_2$$

$$\text{Moles}\ O_2 = (1500\ \cancel{\text{g}\ O_2})\left(\frac{1\ \text{mol}\ O_2}{32.0\ \cancel{\text{g}\ O_2}}\right) = 47\ \text{mol}\ O_2$$

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Sample Exercise 3.19 Calculating the Amount of Product Formed from a Limiting Reactant

Solution (continued)

Thus, there are more moles of H_2 than O_2. The coefficients in the balanced equation indicate, however, that the reaction requires 2 moles of H_2 for every 1 mole of O_2. Therefore, to completely react all the O_2, we would need $2 \times 47 = 94$ moles of H_2. Since there are only 75 moles of H_2, H_2 is the limiting reagent. We therefore use the quantity of H_2 to calculate the quantity of product formed. We can begin this calculation with the grams of H_2, but we can save a step by starting with the moles of H_2 that were calculated previously in the exercise:

$$\text{Grams } H_2O = (75\ \cancel{\text{moles } H_2})\left(\frac{2\ \cancel{\text{mol } H_2O}}{2\ \cancel{\text{mol } H_2}}\right)\left(\frac{18.0\text{ g } H_2O}{1\ \cancel{\text{mol } H_2O}}\right)$$
$$= 1400\text{ g } H_2O \text{ (to two significant figures)}$$

Check The magnitude of the answer seems reasonable. The units are correct, and the number of significant figures (two) corresponds to those in the numbers of grams of the starting materials.

Comment The quantity of the limiting reagent, H_2, can also be used to determine the quantity of O_2 used (37.5 mol = 1200 g). The number of grams of the excess oxygen remaining at the end of the reaction equals the starting amount minus the amount consumed in the reaction, 1500 g – 1200 g = 300 g.

Practice Exercise

A strip of zinc metal with a mass of 2.00 g is placed in an aqueous solution containing 2.50 g of silver nitrate, causing the following reaction to occur:

$$Zn(s) + 2\ AgNO_3(aq) \rightarrow 2\ Ag(s) + Zn(NO_3)_2(aq)$$

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Sample Exercise 3.19 Calculating the Amount of Product Formed from a Limiting Reactant

Practice Exercise (continued)

(a) Which reactant is limiting? **(b)** How many grams of Ag will form? **(c)** How many grams of $Zn(NO_3)_2$ will form? **(d)** How many grams of the excess reactant will be left at the end of the reaction?
Answers: **(a)** $AgNO_3$, **(b)** 1.59 g, **(c)** 1.39 g, **(d)** 1.52 g Zn

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Sample Exercise 3.20 Calculating the Theoretical Yield and the Percent Yield for a Reaction

Adipic acid, $H_2C_6H_8O_4$, is used to produce nylon. The acid is made commercially by a controlled reaction between cyclohexane (C_6H_{12}) and O_2:

$$2\ C_6H_{12}(l) + 5\ O_2(g) \rightarrow 2\ H_2C_6H_8O_4(l) + 2\ H_2O(g)$$

(a) Assume that you carry out this reaction starting with 25.0 g of cyclohexane and that cyclohexane is the limiting reactant. What is the theoretical yield of adipic acid?
(b) If you obtain 33.5 g of adipic acid from your reaction, what is the percent yield of adipic acid?

Solution

Analyze We are given a chemical equation and the quantity of the limiting reactant (25.0 g of C_6H_{12}). We are asked first to calculate the theoretical yield of a product ($H_2C_6H_8O_4$) and then to calculate its percent yield if only 33.5 g of the substance is actually obtained.
Plan (a) The theoretical yield, which is the calculated quantity of adipic acid formed in the reaction, can be calculated using the following sequence of conversions:

$$g\ C_6H_{12} \rightarrow mol\ C_6H_{12} \rightarrow mol\ H_2C_6H_8O_4 \rightarrow g\ H_2C_6H_8O_4$$

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Sample Exercise 3.20 Calculating the Theoretical Yield and the Percent Yield for a Reaction

Solution (continued)

(b) The percent yield is calculated by comparing the actual yield (33.5 g) to the theoretical yield using Equation 3.14.

Solve

$$\text{(a)}\quad \text{Grams } H_2C_6H_8O_4 = (25.0\ \cancel{g\ C_6H_{12}})\left(\frac{1\ \cancel{mol\ C_6H_{12}}}{84.0\ \cancel{g\ C_6H_{12}}}\right)\left(\frac{2\ \cancel{mol\ H_2C_6H_8O_4}}{2\ \cancel{mol\ C_6H_{12}}}\right)\left(\frac{146.0\ g\ H_2C_6H_8O_4}{1\ \cancel{mol\ H_2C_6H_8O_4}}\right)$$

$$= 43.5\ g\ H_2C_6H_8O_4$$

$$\text{(b)}\quad \text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{33.5\ g}{43.5\ g} \times 100\% = 77.0\%$$

Check Our answer in (a) has the appropriate magnitude, units, and significant figures. In (b) the answer is less than 100% as necessary.

Practice Exercise

Imagine that you are working on ways to improve the process by which iron ore containing Fe_2O_3 is converted into iron. In your tests you carry out the following reaction on a small scale:

$$Fe_2O_3(s) + 3\ CO(g) \rightarrow 2\ Fe(s) + 3\ CO_2(g)$$

(a) If you start with 150 g of Fe_2O_3 as the limiting reagent, what is the theoretical yield of Fe? **(b)** If the actual yield of Fe in your test was 87.9 g, what was the percent yield?
Answers: **(a)** 105 g Fe, **(b)** 83.7%

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