CHAPTER 7

Quantities in Chemical Reactions

Chapter Preview

7.1 Stoichiometry

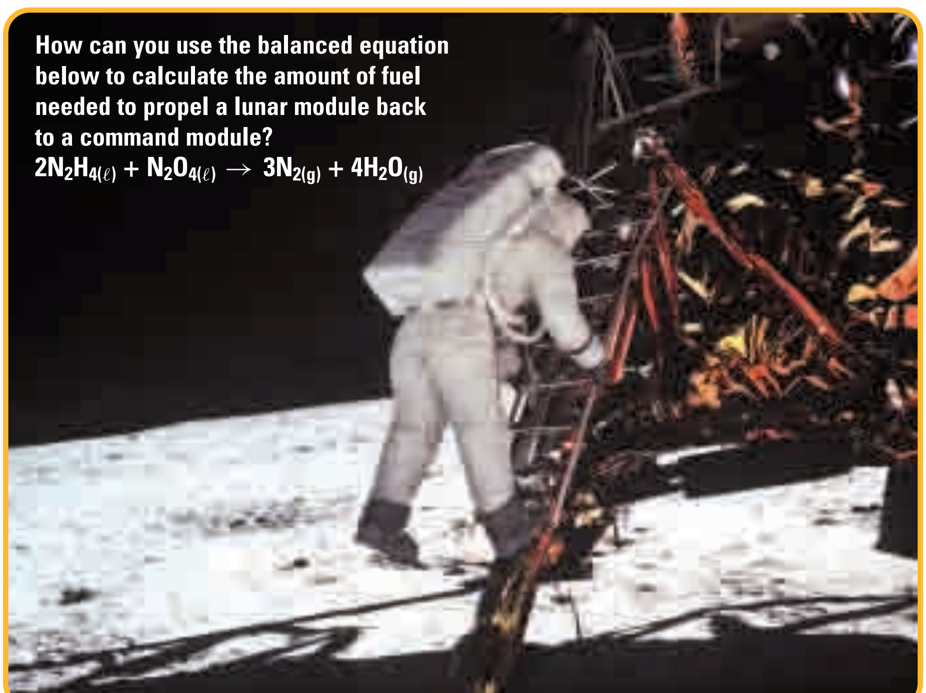
7.2 The Limiting Reactant

7.3 Percentage Yield

A spacecraft, such as the space shuttle on the left, requires a huge amount of fuel to supply the thrust needed to launch it into orbit. Engineers work very hard to minimize the launch mass of a spacecraft because each kilogram requires additional fuel. As well, each kilogram costs thousands of dollars to launch.

In 1969, the *Apollo 11* space mission was the first to land astronauts on the Moon. The engineers on the project faced a challenge when deciding on a fuel for the lunar module. The lunar module took the astronauts from the Moon, back to the command module that was orbiting the Moon. The engineers chose a fuel consisting of hydrazine, N_2H_4, and dinitrogen tetroxide, N_2O_4. These compounds, when mixed, reacted instantaneously and produced the energy needed to launch the lunar module from the Moon.

How do engineers know how much of each reactant they need for a chemical reaction? In this chapter, you will use the concept of the mole to calculate the amounts of reactants that are needed to produce given amounts of products. You will learn how to predict the amounts of products that will be produced in a chemical reaction. You will also learn how to apply this knowledge to any chemical reaction for which you know the balanced chemical equation. Finally, you will learn how calculated amounts deviate from the amounts in real-life situations.



How can you use the balanced equation below to calculate the amount of fuel needed to propel a lunar module back to a command module?

$$2N_2H_{4(\ell)} + N_2O_{4(\ell)} \rightarrow 3N_{2(g)} + 4H_2O_{(g)}$$

Concepts and Skills You Will Need

Before you begin this chapter, review the following concepts and skills:

- balancing chemical equations (Chapter 4, section 4.1)
- understanding the Avogadro constant and the mole (Chapter 5, section 5.2)
- explaining the relationship between the mole and molar mass (Chapter 5, section 5.3)
- solving problems involving number of moles, number of particles, and mass (Chapter 5, section 5.3)

7.1 Stoichiometry

Section Preview/ Specific Expectations

In this section, you will

- **explain** quantitative relationships in a chemical equation, in moles, grams, atoms, or molecules
- **perform** laboratory experiments to determine the meaning of the coefficients in a balanced chemical equation
- **calculate**, for any given reactant or product in a chemical equation, the corresponding mass or quantity of any other reactant or product
- **demonstrate** an awareness of the importance of quantitative chemical relationships in the home or in industry
- **communicate** your understanding of the following terms: *mole ratios, stoichiometry*

Balanced chemical equations are essential for making calculations related to chemical reactions. To understand why, consider the following analogy.

Imagine that you are making salads. You need one head of lettuce, two cucumbers, and five radishes for each salad. Figure 7.1 shows how you can express this as an equation.

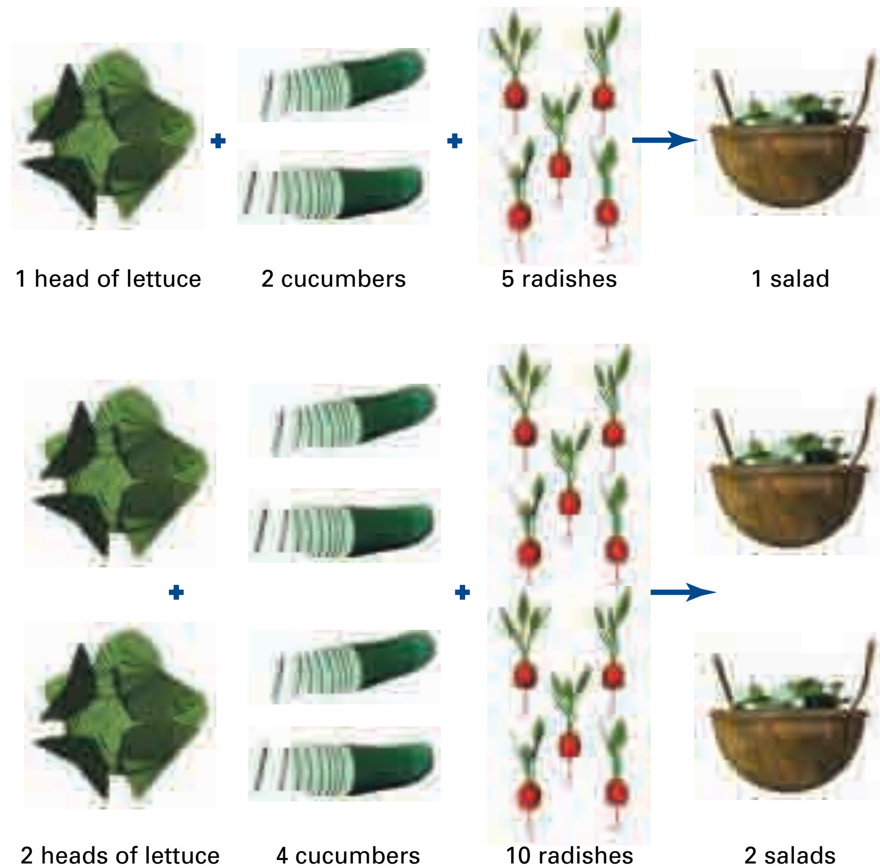


Figure 7.1 **A salad analogy showing how equations can be multiplied**

Now imagine that you are making two salads. How much of each ingredient do you need? You need twice the amount that you used to make one salad, as shown in Figure 7.1.

How many salads can you make if you have three heads of lettuce, six cucumbers, and 15 radishes? According to the salad equation, you can make three salads.

You can get the same kind of information from a balanced chemical equation. In Chapter 4, you learned how to classify chemical reactions and balance the chemical equations that describe them. In Chapters 5 and 6, you learned how chemists relate the number of particles in a substance to the amount of the substance in moles and grams. In this section, you will use your knowledge to interpret the information in a chemical equation, in terms of particles, moles, and mass. Try the following Express Lab to explore the molar relationships between products and reactants.

ExpressLab Mole Relationships in a Chemical Reaction

The following balanced equation shows the reaction between sodium hydrogen carbonate, $NaHCO_3$, and hydrochloric acid, HCl.

$$NaHCO_{3(s)} + HCl_{(aq)} \rightarrow CO_{2(g)} + H_2O_{(\ell)} + NaCl_{(aq)}$$

In this Express Lab, you will determine the mole relationships between the products and reactants in the reaction. Then you will compare the mole relationships with the balanced chemical equation.

Safety Precautions

Be careful when using concentrated hydrochloric acid. It burns skin and clothing. Do not inhale its vapour.

Procedure

1. Obtain a sample of sodium hydrogen carbonate that is approximately 1.0 g.
2. Place a 24-well microplate on a balance. Measure and record its mass.
3. Place all the sodium hydrogen carbonate in well A4 of the microplate. Measure and record the mass of the microplate and sample.
4. Fill a thin-stem pipette with 8 mol/L hydrochloric acid solution.
5. Wipe the outside of the pipette. Stand it, stem up, in well A3.
6. Measure and record the total mass of the microplate and sample.
7. Add the hydrochloric acid from the pipette to the sodium hydrogen carbonate in well A4. Allow the gas to escape after each drop.
8. Continue to add the hydrochloric acid until all the sodium hydrogen carbonate has dissolved and the solution produces no more bubbles.
9. Return the pipette, stem up, to well A3. Again find the total mass of the microplate and samples.
10. Dispose of the reacted chemicals as directed by your teacher.

Analysis

1. Calculate the number of moles of sodium hydrogen carbonate used.
2. Find the difference between the total mass of the microplate and samples before and after the reaction. This difference represents the mass of carbon dioxide gas produced.
3. Calculate the number of moles of carbon dioxide produced.
4. Express your answers to questions 1 and 3 as a mole ratio of mol $NaHCO_3$:mol CO_2.
5. According to the balanced equation, how many molecules of sodium hydrogen carbonate react to form one molecule of carbon dioxide?
   (a) Express your answer as a ratio.
   (b) Compare this ratio to your mole ratio in question 4.
6. How many moles of carbon dioxide do you think would be formed from 4.0 mol of sodium hydrogen carbonate?

You can use your understanding of the relationship between moles and number of particles to see how chemical equations communicate information about how many moles of products and reactants are involved in a reaction.

Particle Relationships in a Balanced Chemical Equation

As you learned in Chapter 4, the coefficients in front of compounds and elements in chemical equations tell you how many atoms and molecules participate in a reaction. A chemical equation can tell you much more, however. Consider, for example, the equation that describes the production of ammonia. Ammonia is an important industrial chemical. Several of its uses are shown in Figure 7.2 on the following page.

Figure 7.2 **Ammonia can be applied directly to the soil as a fertilizer. An aqueous (water) solution of ammonia can be used as a household cleaner.**

Ammonia can be prepared industrially from its elements, using a process called the Haber Process. The Haber Process is based on the balanced chemical equation below.

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

This equation tells you that one molecule of nitrogen gas reacts with three molecules of hydrogen gas to form two molecules of ammonia gas.

As you can see in Figure 7.3, there is the same number of each type of atom on both sides of the equation.

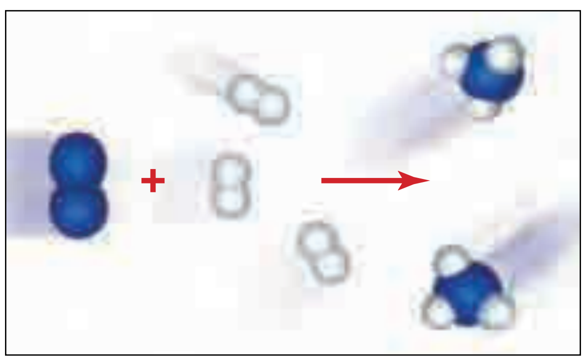


Figure 7.3 **The reaction of nitrogen gas with hydrogen gas.**

You can use a ratio to express the numbers of atoms in the equation, as follows:

$$1 \text{ molecule } N_2 : 3 \text{ molecules } H_2 : 2 \text{ molecules } NH_3$$

What happens if you multiply the ratio by 2? You get

$$2 \text{ molecules } N_2 : 6 \text{ molecules } H_2 : 4 \text{ molecules } NH_3$$

This means that two molecules of nitrogen gas react with six molecules of hydrogen gas to produce four molecules of ammonia gas. Multiplying the original ratio by one dozen gives the following relationship:

$$1 \text{ dozen molecules } N_2 : 3 \text{ dozen molecules } H_2 : 2 \text{ dozen molecules } NH_3$$

Suppose that you want to produce 20 molecules of ammonia. How many molecules of nitrogen do you need? You know that you need one molecule of nitrogen for every two molecules of ammonia produced. In other words, the number of molecules of nitrogen that you need is one half the number of molecules of ammonia that you want to produce.

$$20 \cancel{\text{ molecules } NH_3} \times \frac{1 \text{ molecule } N_2}{2 \cancel{\text{ molecules } NH_3}} = 10 \text{ molecules } N_2$$

Try the following problems to practise working with ratios in balanced chemical equations.

Practice Problems

1. Consider the following reaction. 1 2

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(\ell)}$$

2 2:1:2

(a) Write the ratio of H_2 molecules: O_2 molecules: H_2O molecules.

(b) How many molecules of O_2 are required to react with 100 molecules of H_2, according to your ratio in part (a)? 50

(c) How many molecules of water are formed when 2478 molecules of O_2 react with H_2? $2478 \times 2 = ?$

(d) How many molecules of H_2 are required to react completely with 6.02×10^{23} molecules of O_2? 1 mol O_2 , 2 mol H_2 .

2. Iron reacts with chlorine gas to form iron(III) chloride, $FeCl_3$.

$$2Fe_{(s)} + 3Cl_{2(g)} \rightarrow 2FeCl_{3(s)}$$

(a) How many atoms of Fe are needed to react with three molecules of Cl_2?

(b) How many molecules of $FeCl_3$ are formed when 150 atoms of Fe react with sufficient Cl_2?

(c) How many Cl_2 molecules are needed to react with 1.204×10^{24} atoms of Fe?

(d) How many molecules of $FeCl_3$ are formed when 1.806×10^{24} molecules of Cl_2 react with sufficient Fe?

3. Consider the following reaction. 1 2 盐酸.

$$Ca(OH)_{2(aq)} + 2HCl_{(aq)} \rightarrow CaCl_2 + 2H_2O_{(\ell)}$$

1 1

(a) How many formula units of calcium chloride, $CaCl_2$, would be produced by 6.7×10^{25} molecules of hydrochloric acid, HCl?

(b) How many molecules of water would be produced in the reaction in part (a)?

$$\frac{HCl}{CaCl_2} = \frac{2}{1} = \frac{x}{6.7 \times 10^{25}}$$

Mole Relationships in Chemical Equations

Until now, you have assumed that the coefficients in a chemical equation represent particles. They can, however, also represent moles. Consider the following ratio to find out why.

$$1 \text{ molecule } N_2 : 3 \text{ molecules } H_2 : 2 \text{ molecules } NH_3$$

You can multiply the above ratio by the Avogadro constant to obtain

$1 \times N_A$ molecules N_2 : $3 \times N_A$ molecules H_2 : $2 \times N_A$ molecules NH_3

This is the same as

$$1 \text{ mol } N_2 : 3 \text{ mol } H_2 : 2 \text{ mol } NH_3$$

So the chemical equation $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$ also means that 1 mol of nitrogen molecules reacts with 3 mol of hydrogen molecules to form 2 mol of ammonia molecules. The relationships between moles in a balanced chemical equation are called **mole ratios**. For example, the mole ratio of nitrogen to hydrogen in the equation above is 1 mol N_2:3 mol H_2. The mole ratio of hydrogen to ammonia is 3 mol H_2:2 mol NH_3.

You can manipulate mole ratios in the same way that you can manipulate ratios involving molecules. For example, suppose that you want to know how many moles of ammonia are produced by 2.8 mol of hydrogen. You know that you can obtain 2 mol of ammonia for every 3 mol of hydrogen. Therefore, you multiply the number of moles of hydrogen by the mole ratio of ammonia to hydrogen. Another way to think about this is to equate the known mole ratio of hydrogen to ammonia to the unknown mole ratio of hydrogen to ammonia and solve for the unknown.

$$\underset{\textbf{unknown ratio}}{\frac{n \text{ mol } NH_3}{2.8 \text{ mol } H_2}} = \underset{\textbf{known ratio}}{\frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2}}$$

$$(\cancel{2.8 \text{ mol } H_2})\frac{n \text{ mol } NH_3}{\cancel{2.8 \text{ mol } H_2}} = (2.8\ \cancel{\text{mol } H_2})\frac{2 \text{ mol } NH_3}{3\ \cancel{\text{mol } H_2}}$$

$$n \text{ mol } NH_3 = 1.9 \text{ mol } NH_3$$

Try the following Practice Problems to work with mole ratios.

CHEM FACT

Because the coefficients of a balanced chemical equation can represent moles, it is acceptable to use fractions in an equation. For example, you can write the equation

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(\ell)}$

as

$H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(\ell)}$

Half an oxygen molecule is an oxygen atom, which does not accurately reflect the reaction. Half a mole of oxygen molecules, however, makes sense.

Practice Problems

4. Aluminum bromide can be prepared by reacting small pieces of aluminum foil with liquid bromine at room temperature. The reaction is accompanied by flashes of red light.

$$2Al_{(s)} + 3Br_{2(\ell)} \rightarrow 2AlBr_{3(s)}$$

How many moles of Br_2 are needed to produce 5 mol of $AlBr_3$, if sufficient Al is present?

5. Hydrogen cyanide gas, $HCN_{(g)}$, is used to prepare clear, hard plastics, such as Plexiglas™. Hydrogen cyanide is formed by reacting ammonia, NH_3, with oxygen and methane, CH_4.

$$2NH_{3(g)} + 3O_{2(g)} + 2CH_{4(g)} \rightarrow 2HCN_{(g)} + 6H_2O_{(g)}$$

(a) How many moles of O_2 are needed to react with 1.2 mol of NH_3?

(b) How many moles of H_2O can be expected from the reaction of 12.5 mol of CH_4? Assume that sufficient NH_3 and O_2 are present.

6. Ethane gas, C_2H_6, is present in small amounts in natural gas. It undergoes complete combustion to produce carbon dioxide and water.

$$2C_2H_{6(g)} + 7O_{2(g)} \rightarrow 4CO_{2(g)} + 6H_2O_{(g)}$$

(a) How many moles of O_2 are required to react with 13.9 mol of C_2H_6?

(b) How many moles of H_2O would be produced by 1.40 mol of O_2 and sufficient ethane?

7. Magnesium nitride reacts with water to produce magnesium hydroxide and ammonia gas, NH_3 according to the balanced chemical equation

$$Mg_3N_{2(s)} + 6H_2O_{(\ell)} \rightarrow 3Mg(OH)_{2(s)} + 2NH_{3(g)}$$

a) How many molecules of water are required to react with 2.3 mol Mg_3N_2?

b) How many molecules of $Mg(OH)_2$ will be expected in part (a)?

Different Ratios of Reactants

The relative amounts of reactants are important. Different mole ratios of the same reactants can produce different products. For example, carbon can combine with oxygen in two different ratios, forming either carbon monoxide or carbon dioxide. In the following reaction, the mole ratio of carbon to oxygen is 2 mol C:1 mol O_2.

$$2C_{(s)} + O_{2(g)} \rightarrow 2CO_{(g)}$$

In the next reaction, the mole ratio of carbon to oxygen is 1 mol C:1 mol O_2.

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$$

Thus, carbon dioxide forms if carbon and oxygen are present in a mole ratio of about 1 mol C:1 mol O_2. Carbon dioxide is a product of cellular respiration in animals and humans, and it is a starting material for photosynthesis. It is also one of the products of the complete combustion of a hydrocarbon fuel.

If there is a relative shortage of oxygen, however, and the mole ratio of carbon to oxygen is closer to 2 mol C:1 mol O, carbon monoxide forms. Carbon monoxide is colourless, tasteless, and odourless. It is a highly poisonous gas, that is responsible for the deaths of hundreds of people in Canada and the United States every year. Carbon monoxide can escape from any fuel-burning appliance: furnace, water heater, fireplace, wood stove, or space heater. If you have one of these appliances in your home, make sure that it has a good supply of oxygen to avoid the formation of carbon monoxide.

There are many reactions in which different mole ratios of the reactants result in different products. The following Sample Problem will help you understand how to work with these reactions.

Technology LINK

In many areas, it is mandatory for every home to have a carbon monoxide detector, like the one shown below. If you do not have a carbon monoxide detector in your home, you can buy one at a hardware store for a modest price. It could end up saving your life.

A carbon monoxide detector emits a sound when the level of carbon monoxide exceeds a certain limit. Find out how a carbon monoxide detector works, and where it should be placed. Present your findings as a public service announcement.

Sample Problem

Mole Ratios of Reactants

Problem

Vanadium can form several different compounds with oxygen, including V_2O_5, VO_2, and V_2O_3. Determine the number of moles of oxygen that are needed to react with 0.56 mol of vanadium to form divanadium pentoxide, V_2O_5.

What Is Required?

You need to find the number of moles of oxygen that are needed to react with 0.56 mol of vanadium to form divanadium pentoxide.

What Is Given?

Reactant: vanadium, V $\rightarrow$ 0.56 mol
Reactant: oxygen, O_2
Product: divanadium pentoxide, V_2O_5

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Plan Your Strategy

Write a balanced chemical equation for the formation of vanadium(V) oxide. Use the known mole ratio of vanadium to oxygen to calculate the unknown amount of oxygen.

Act on Your Strategy

The balanced equation is $4V_{(s)} + 5O_{2(g)} \rightarrow 2V_2O_{5(s)}$

To determine the number of moles of oxygen required, equate the known ratio of oxygen to vanadium from the balanced equation to the unknown ratio from the question.

$$\underset{\textbf{unknown ratio}}{\frac{n \text{ mol } O_2}{0.56 \text{ mol V}}} = \underset{\textbf{known ratio}}{\frac{5 \text{ mol } O_2}{4 \text{ mol V}}}$$

Multiply both sides of the equation by 0.56 mol V.

$$(\cancel{0.56 \text{ mol V}}) \frac{n \text{ mol } O_2}{\cancel{0.56 \text{ mol V}}} = (0.56 \text{ mol V}) \frac{5 \text{ mol } O_2}{4 \text{ mol V}}$$

$$n \text{ mol } O_2 = (0.56 \cancel{\text{ mol V}}) \frac{5 \text{ mol } O_2}{4 \cancel{\text{ mol V}}}$$

$$= 0.70 \text{ mol } O_2$$

Check Your Solution

The units are correct. The mole ratio of vanadium to oxygen is 4 mol V:5 mol O_2. Multiply 0.70 mol by 4/5, and you get 0.56 mol. The answer is therefore reasonable.

Practice Problems

$V + O_2 \rightarrow VO_2$

7.47 = 7.47 = 7.47 mol

8. Refer to the Sample Problem above.

(a) How many moles of V are needed to produce 7.47 mol of VO_2? Assume that sufficient O_2 is present.

(1) 5.39 × 4

(b) How many moles of V are needed to react with 5.39 mol of O_2 to produce V_2O_3? $4V + 3O_2 \longrightarrow 2V_2O_3$

$\frac{O_2}{V} = \frac{5.39}{x} = \frac{3}{4}$

$x = \frac{5.39 \times 4}{3}$

9. Nitrogen, N_2, can combine with oxygen, O_2, to form several different oxides of nitrogen. These oxides include NO_2, NO, and N_2O.

(a) How many moles of O_2 are required to react with 9.35×10^{-2} moles of N_2 to form N_2O?

(b) How many moles of O_2 are required to react with 9.35×10^{-2} moles of N_2 to form NO_2?

10. When heated in a nickel vessel to 400°C, xenon can be made to react with fluorine to produce colourless crystals of xenon tetrafluoride.

a) How many moles of fluorine gas, F_2, would be required to react with 3.54×10^{-1} mol of xenon?

b) Under somewhat similar reaction conditions, xenon hexafluoride can also be obtained. How many moles of fluorine would be required to react with the amount of xenon given in part (a) to produce xenon hexafluoride?

CHECKPOINT

Do you think that xenon could be made to react with bromine or iodine under the same conditions outlined in Practice Problem 10? Explain why or why not, using your understanding of periodic trends.

Mass Relationships in Chemical Equations

As you have learned, the coefficients in a balanced chemical equation represent moles as well as particles. Therefore, you can use the molar masses of reactants and products to determine the mass ratios for a reaction. For example, consider the equation for the formation of ammonia:

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

You can find the mass of each substance using the equation $m = M \times n$ as follows:

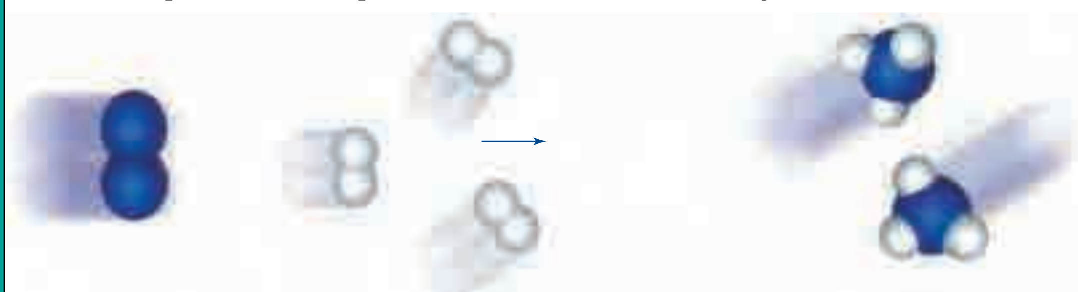
$$1 \text{ mol } N_2 \times 28.0 \text{ g/mol } N_2 = 28.0 \text{ g } N_2$$
$$3 \text{ mol } H_2 \times 2.02 \text{ g/mol } H_2 = 6.1 \text{ g } H_2$$
$$2 \text{ mol } NH_3 \times 17.0 \text{ g/mol } NH_3 = 34.1 \text{ g } NH_3$$

In Table 7.1, you can see how particles, moles, and mass are related in a chemical equation. Notice that the mass of the product is equal to the total mass of the reactants. This confirms the law of conservation of mass.

CHECKPOINT

Refer back to Chapter 5. Calculate the molar mass of N_2, H_2, and NH_3.

Table 7.1 What a Balanced Chemical Equation Tells You

Balanced equation	$N_{2(g)} + 3H_{2(g)}$	→	$2NH_{3(g)}$
Number of particles (molecules)	1 molecule N_2 + 3 molecules H_2 	→	2 molecules NH_3
Amount (mol)	1 mol N_2 + 3 mol H_2	→	2 mol NH_3
Mass (g)	28.0 g N_2 + 6.1 g H_2	→	34.1 g NH_3
Total mass (g)	34.1 g reactants	→	34.1 g product

Stoichiometric Mass Calculations

You now know what a balanced chemical equation tells you in terms of number of particles, number of moles, and mass of products and reactants. How do you use this information? Because reactants and products are related by a fixed ratio, if you know the number of moles of one substance, the balanced equation tells you the number of moles of all the other substances. In Chapters 5 and 6, you learned how to convert between particles, moles, and mass. Therefore, *if you know the amount of one substance in a chemical reaction (in particles, moles, or mass), you can calculate the amount of any other substance in the reaction (in particles, moles, or mass), using the information in the balanced chemical equation.*

You can see that a balanced chemical equation is a powerful tool. It allows chemists to predict the amount of products that will result from a reaction involving a known amount of reactants. As well, chemists can use a balanced equation to calculate the amount of reactants they will need to produce a desired amount of products. They can also use it to predict the amount of one reactant they will need to react completely with another reactant.

Language LINK

The word "stoichiometry" is derived from two Greek words: *stoikheion*, meaning "element," and *metron*, meaning "to measure." What other words might be derived from the Greek word *metron*?

Stoichiometry is the study of the relative quantities of reactants and products in chemical reactions. Stoichiometric calculations are used for many purposes. One purpose is determining how much of a reactant is needed to carry out a reaction. This kind of knowledge is useful for any chemical reaction, and it can even be a matter of life or death.

In a spacecraft, for example, carbon dioxide is produced as the astronauts breathe. To maintain a low level of carbon dioxide, air in the cabin is passed continuously through canisters of lithium hydroxide granules. The carbon dioxide reacts with the lithium hydroxide in the following way:

$$CO_{2(g)} + 2LiOH_{(s)} \rightarrow Li_2CO_{3(s)} + H_2O_{(g)}$$

The canisters are changed periodically as the lithium hydroxide reacts. Engineers must calculate the amount of lithium hydroxide needed to ensure that the carbon dioxide level is safe. As you learned earlier, every kilogram counts in space travel. Therefore, a spacecraft cannot carry much more than the minimum amount.

Figure 7.4 **A spacecraft is a closed system. All chemical reactions must be taken into account when engineers design systems to keep the air breathable.**

History LINK

The concept of stoichiometry was first described in 1792 by the German scientist Jeremias Benjamin Richter (1762–1807). He stated that "stoichiometry is the science of measuring the quantitative proportions or mass ratios in which chemical elements stand to one another." Can you think of another reason why Richter was famous?

To determine how much lithium hydroxide is needed, engineers need to ask and answer two important questions:

- How much carbon dioxide is produced per astronaut each day?
- How much lithium hydroxide is needed per kilogram of carbon dioxide?

Engineers can answer the first question by experimenting. To answer the second question, they need to do stoichiometric calculations. Examine the following Sample Problems to see how these calculations would be done.

Sample Problem

Mass to Mass Calculations for Reactants

Problem

Carbon dioxide that is produced by astronauts can be removed with lithium hydroxide. The reaction produces lithium carbonate and water. An astronaut produces an average of 1.00×10^3 g of carbon dioxide each day. What mass of lithium hydroxide should engineers put on board a spacecraft, per astronaut, for each day?

What Is Required?

You need to find the mass of lithium hydroxide that is needed to react with 1.00×10^3 g of carbon dioxide.

What Is Given?

Reactant: carbon dioxide, $CO_2 \rightarrow 1.00 \times 10^3$ g
Reactant: lithium hydroxide, LiOH
Product: lithium carbonate, Li_2CO_3
Product: water, H_2O

Plan Your Strategy

Step 1 Write a balanced chemical equation.

Step 2 Convert the given mass of carbon dioxide to the number of moles of carbon dioxide.

Step 3 Calculate the number of moles of lithium hydroxide based on the mole ratio of lithium hydroxide to carbon dioxide.

Step 4 Convert the number of moles of lithium hydroxide to grams.

Act on Your Strategy

The balanced chemical equation is

1 $$CO_{2(g)} + 2LiOH_{(s)} \longrightarrow Li_2CO_{3(s)} + H_2O_{(g)}$$

22.7 mol → 45.4 mol

3 unknown ratio = known ratio

$$\frac{n \text{ mol LiOH}}{22.7 \text{ mol } CO_2} = \frac{2 \text{ mol LiOH}}{1 \text{ mol } CO_2}$$

$$(\cancel{22.7 \text{ mol } CO_2})\frac{n \text{ mol LiOH}}{\cancel{22.7 \text{ mol } CO_2}} = \frac{2 \text{ mol LiOH}}{1 \cancel{\text{ mol } CO_2}}(22.7 \cancel{\text{ mol } CO_2})$$

$$n \text{ mol LiOH} = 45.4 \text{ mol LiOH}$$

2 $$\frac{1.00 \times 10^3 \cancel{\text{g}}\, CO_2}{44.0 \cancel{\text{g}}\text{ mol}} = 22.7 \text{ mol } CO_2$$

4 $$45.4 \cancel{\text{mol}} \text{ LiOH} \times 23.9 \text{ g}/\cancel{\text{mol}} \text{ LiOH} = 1.09 \times 10^3 \text{ g LiOH}$$

1.00×10^3 g CO_2 → 1.09×10^3 g LiOH

Continued ...

CHEM FACT

The Group 18 elements in the periodic table are currently called the noble gases. In the past, however, they were referred to as the inert gases. They were believed to be totally unreactive. Scientists have found that this is not true. Some of them can be made to react with reactive elements, such as fluorine, under the proper conditions. In 1962, the synthesis of the first compound that contained a noble gas was reported. Since then, a number of noble gas compounds have been prepared, mostly from xenon. A few compounds of krypton, radon, and argon have also been prepared.

In the early 1960s, Neil Bartlett, of the University of British Columbia, synthesized the first compound that contained a noble gas.

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Therefore, 1.09×10^3 g LiOH are required.

Check Your Solution

The units are correct. Lithium hydroxide has a molar mass that is about half of carbon dioxide's molar mass, but there are twice as many moles of lithium hydroxide. Therefore it makes sense that the mass of lithium hydroxide required is about the same as the mass of carbon dioxide produced.

Practice Problems

11. Ammonium sulfate, $(NH_4)_2SO_4$, is used as a source of nitrogen in some fertilizers. It reacts with sodium hydroxide to produce sodium sulfate, water and ammonia.

$$(NH_4)_2SO_{4(s)} + 2NaOH_{(aq)} \rightarrow Na_2SO_{4(aq)} + 2NH_{3(g)} + 2H_2O_{(\ell)}$$

What mass of sodium hydroxide is required to react completely with 15.4 g of $(NH_4)_2SO_4$?

12. Iron(III) oxide, also known as rust, can be removed from iron by reacting it with hydrochloric acid to produce iron(III) chloride and water.

$$Fe_2O_{3(s)} + 6HCl_{(aq)} \rightarrow 2FeCl_{3(aq)} + 3H_2O_{(\ell)}$$

What mass of hydrogen chloride is required to react with 1.00×10^2 g of rust?

13. Iron reacts slowly with hydrochloric acid to produce iron(II) chloride and hydrogen gas. 0.0637 × 2 × (36.458) =

0.0637   1   2   1   1

$$Fe_{(s)} + 2HCl_{(aq)} \rightarrow FeCl_{2(aq)} + H_{2(g)}$$

① 3.56 × 2 g.

What mass of HCl is required to react with 3.56 g of iron?

$\frac{3.56}{55.845}$ = mol 0.0637

14. Dinitrogen pentoxide is a white solid. When heated it decomposes to produce nitrogen dioxide and oxygen.

$$2N_2O_{5(s)} \rightarrow 4NO_{2(g)} + O_{2(g)}$$

How many grams of oxygen gas will be produced in this reaction when 2.34 g of NO_2 are made?

Sample Problem

Mass to Mass Calculations for Products and Reactants

Problem

In the Chapter 7 opener, you learned that a fuel mixture consisting of hydrazine, N_2H_4, and dinitrogen tetroxide, N_2O_4, was used to launch a lunar module. These two compounds react to form nitrogen gas and water vapour. If 50.0 g of hydrazine reacts with sufficient dinitrogen tetroxide, what mass of nitrogen gas is formed?

Continued ...

Continued ...
FROM PAGE 244

What Is Required?

You need to find the mass of nitrogen gas that is formed from 50.0 g of hydrazine.

What Is Given?

Reactant: hydrazine, $N_2H_4 \rightarrow 150.0$ g
Reactant: dinitrogen tetroxide, N_2O_4
Product: nitrogen, N_2
Product: water, H_2O

Plan Your Strategy

Step 1 Write a balanced chemical equation.

Step 2 Convert the mass of hydrazine to the number of moles of hydrazine.

Step 3 Calculate the number of moles of nitrogen, using the mole ratio of hydrazine to nitrogen.

Step 4 Convert the number of moles of nitrogen to grams.

Electronic Learning Partner

Go to your Chemistry 11 Electronic Learning Partner for a video clip showing an experiment that uses stoichiometry.

Act on Your Strategy

The balanced chemical equation is

1

$$2N_2H_{4(\ell)} + N_2O_{4(\ell)} \longrightarrow 3N_{2(g)} + 4H_2O_{(g)}$$

4.679 mol → 7.019 mol

3

unknown ratio = known ratio

$$\frac{n \text{ mol } N_2}{4.680 \text{ mol } N_2H_4} = \frac{3 \text{ mol } N_2}{2 \text{ mol } N_2H_4}$$

$$(4.679 \cancel{\text{ mol } N_2H_4})\frac{n \text{ mol } N_2}{4.679 \cancel{\text{ mol } N_2H_4}} = \frac{3 \text{ mol } N_2}{2 \cancel{\text{ mol } N_2H_4}}(4.679 \cancel{\text{ mol } N_2H_4})$$

$$= 7.019 \text{ mol } N_2$$

2

$$\frac{150.0 \cancel{\text{ g}} N_2H_4}{32.06 \cancel{\text{ g}}/\text{mol}} = 4.6 \text{ mol}$$

4

$$7.019 \cancel{\text{ mol } N_2} \times 28.01 \text{ g}/\cancel{\text{mol } N_2} = 196.6 \text{ g}$$

150.0 g N_2H_4 → 196.6 g N_2

Therefore, 196.6 g of nitrogen are formed.

Check Your Solution

The units are correct. Nitrogen has a molar mass that is close to hydrazine's molar mass. Therefore, to estimate the amount of nitrogen from the mass of hydrazine, multiply the mole ratio of nitrogen to hydrazine (3:2) by hydrazine's mass (150 g) to get 225 g, which is close to the calculated answer, 196.6 g. The answer is reasonable.

Continued ...
FROM PAGE 245

Practice Problems

15. Powdered zinc reacts rapidly with powdered sulfur in a highly exothermic reaction.

$$8Zn_{(s)} + S_{8(s)} \rightarrow 8ZnS_{(s)}$$

What mass of zinc sulfide is expected when 32.0 g of S_8 reacts with sufficient zinc?

16. The addition of concentrated hydrochloric acid to manganese(IV) oxide leads to the production of chlorine gas.

$$4HCl_{(aq)} + MnO_{2(g)} \rightarrow MnCl_{2(aq)} + Cl_{2(g)} + 2H_2O_{(\ell)}$$

What mass of chlorine can be obtained when 4.76×10^{-2} g of HCl react with sufficient MnO_2?

17. Aluminum carbide, Al_4C_3, is a yellow powder that reacts with water to produce aluminum hydroxide and methane.

$$Al_4C_{3(s)} + 12H_2O_{(\ell)} \rightarrow 4Al(OH)_{3(s)} + 3CH_{4(g)}$$

What mass of water is required to react completely with 25.0 g of aluminum carbide?

18. Magnesium oxide reacts with phosphoric acid, H_3PO_4, to produce magnesium phosphate and water.

$$3MgO_{(s)} + 2H_3PO_{4(aq)} \rightarrow Mg_3(PO_4)_{2(s)} + 3H_2O_{(\ell)}$$

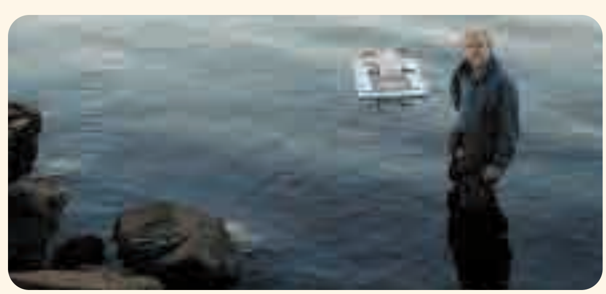
How many grams of magnesium oxide are required to react completely with 33.5 g of phosphoric acid?

Canadians in Chemistry

As a chemist with Environment Canada's Atmospheric Science Division in Dartmouth, Nova Scotia, Dr. Stephen Beauchamp studies toxic chemicals, such as mercury. Loons in Nova Scotia's Kejimkujik National Park are among the living creatures that he studies. Kejimkujik loons have higher blood mercury levels (5 μg Hg/1 g blood) than any other North American loons (2 μg Hg/1 g blood). Mercury is also found in high levels in the fish the loons eat. Mercury causes behavioural problems in the loons. As well, it may affect the loons' reproductive success and immune function.

Bacteria convert environmental mercury into methyl mercury, CH_3Hg. This is the form that is most easily absorbed into living organisms. Beauchamp examines forms and concentrations of mercury in the air, soil, and water.

Mercury emission sources include electrical power generation, manufacturing, and municipal waste incineration. Sources such as these, however, do not totally account for the high mercury levels found in Kejimkujik loons and other area wildlife. Beauchamp is working to discover what other factors are operating so that he will be able to recommend ways to improve the situation.



Dr. Stephen Beauchamp in Halifax Harbour. The flux chamber beside him helps him measure the changing concentrations of mercury in the air and water.

A General Process for Solving Stoichiometric Problems

You have just solved several stoichiometric problems. In these problems, masses of products and reactants were given, and masses were also required for the answers. Chemists usually need to know what mass of reactants they require and what mass of products they can expect. Sometimes, however, a question requires you to work with the number of moles or particles. Use the same process for solving stoichiometric problems, whether you are working with mass, moles, or particles:

Step 1 Write a balanced chemical equation.

Step 2 If you are given the mass or number of particles of a substance, convert it to the number of moles.

Step 3 Calculate the number of moles of the required substance based on the number of moles of the given substance, using the appropriate mole ratio.

Step 4 Convert the number of moles of the required substance to mass or number of particles, as directed by the question.

Examine the following Sample Problem to see how to work with mass and particles.

Sample Problem

Mass and Particle Stoichiometry

Problem

Passing chlorine gas through molten sulfur produces liquid disulfur dichloride. How many molecules of chlorine react to produce 50.0 g of disulfur dichloride?

What Is Required?

You need to determine the number of molecules of chlorine gas that produce 50.0 g of disulfur dichloride.

What Is Given?

Reactant: chlorine, Cl_2
Reactant: sulfur, S
Product: disulfur dichloride, $S_2Cl_2 \rightarrow$ 50.0 g

Plan Your Strategy

Step 1 Write a balanced chemical equation.

Step 2 Convert the given mass of disulfur dichloride to the number of moles.

Step 3 Calculate the number of moles of chlorine gas using the mole ratio of chlorine to disulfur dichloride.

Step 4 Convert the number of moles of chlorine gas to the number of particles of chlorine gas.

Continued ...

Continued ...
FROM PAGE 247

Act on Your Strategy

1 $Cl_{2(g)} + 2S_{(\ell)} \longrightarrow S_2Cl_{2(\ell)}$

0.370 mol ← 0.370 mol

unknown ratio **known ratio**

3
$$\frac{\text{amount } Cl_2}{0.370 \text{ mol } S_2Cl_2} = \frac{1 \text{ mol } Cl_2}{1 \text{ mol } S_2Cl_2}$$

$$(\cancel{0.370 \text{ mol } S_2Cl_2})\frac{\text{amount } Cl_2}{\cancel{0.370 \text{ mol } S_2Cl_2}} = (0.370 \text{ mol } \cancel{S_2Cl_2})\frac{1 \text{ mol } Cl_2}{1 \text{ mol } \cancel{S_2Cl_2}}$$

$$\text{amount } Cl_2 = 0.370 \text{ mol } Cl_2$$

2
$$\frac{50.0 \cancel{\text{g}} S_2Cl_2}{135 \cancel{\text{g}}/\text{mol}} = 0.370 \text{ mol } S_2Cl_2$$

4
$$0.370 \cancel{\text{mol } Cl_2} \times 6.02 \times 10^{23} \frac{\text{molecules } Cl_2}{\cancel{\text{mol } Cl_2}}$$
$$= 2.22 \times 10^{23} \text{ molecules } Cl_2$$

2.22×10^{23} molecules Cl_2 ← 50.0 g S_2Cl_2

Therefore, 2.22×10^{23} molecules of chlorine gas are required.

Check Your Solution

The units are correct. 2.0×10^{23} is about 1/3 of a mole, or 0.33 mol. One-third of a mole of disulfur dichloride has a mass of 45 g, which is close to 50 g. The answer is reasonable.

The thermite reaction generates enough heat to melt the elemental iron that is produced.

Practice Problems

19. Nitrogen gas is produced in an automobile air bag. It is generated by the decomposition of sodium azide, NaN_3.

$$2NaN_{3(s)} \rightarrow 3N_{2(g)} + 2Na_{(s)}$$

(a) To inflate the air bag on the driver's side of a certain car, 80.0 g of N_2 is required. What mass of NaN_3 is needed to produce 80.0 g of N_2?

(b) How many atoms of Na are produced when 80.0 g of N_2 are generated in this reaction?

20. The reaction of iron(III) oxide with powdered aluminum is known as the thermite reaction.

$$2Al_{(s)} + Fe_2O_{3(s)} \rightarrow Al_2O_{3(s)} + 2Fe_{(\ell)}$$

(a) Calculate the mass of aluminum oxide, Al_2O_3, that is produced when 1.42×10^{24} atoms of Al react with Fe_2O_3.

(b) How many formula units of Fe_2O_3 are needed to react with 0.134 g of Al?

Continued ...

Continued ...
FROM PAGE 248

21. The thermal decomposition of ammonium dichromate is an impressive reaction. When heated with a Bunsen burner or propane torch, the orange crystals of ammonium dichromate slowly decompose to green chromium(III) oxide in a volcano-like display. Colourless nitrogen gas and water vapour are also given off.

$$(NH_4)_2Cr_2O_{7(s)} \rightarrow Cr_2O_{3(s)} + N_{2(g)} + 4H_2O_{(g)}$$

(a) Calculate the number of molecules of Cr_2O_3 that is produced from the decomposition of 10.0 g of $(NH_4)_2Cr_2O_7$.

(b) In a different reaction, 16.9 g of N_2 is produced when a sample of $(NH_4)_2Cr_2O_7$ is decomposed. How many water molecules are also produced in this reaction?

(c) How many formula units of $(NH_4)_2Cr_2O_7$ are needed to produce 1.45 g of H_2O?

22. Ammonia gas reacts with oxygen to produce water and nitrogen oxide. This reaction can be catalyzed, or sped up, by Cr_2O_3, produced in the reaction in problem 21.

$$4NH_{3(g)} + 5O_{2(g)} \rightarrow 4NO_{(g)} + 6H_2O_{(\ell)}$$

(a) How many molecules of oxygen are required to react with 34.0 g of ammonia?

(b) What mass of nitrogen oxide is expected from the reaction of 8.95×10^{24} molecules of oxygen with sufficient ammonia?

Section Wrap-up

You have learned how to do stoichiometric calculations, using balanced chemical equations to find amounts of reactants and products. In these calculations, you assumed that the reactants and products occurred in the exact molar ratios shown by the chemical equation. In real life, however, reactants are often not present in these exact ratios. Similarly, the amount of product that is predicted by stoichiometry is not always produced. In the next two sections, you will learn how chemists deal with these challenges.

Section Review

1 K/U Why is a balanced chemical equation needed to solve stoichiometric calculations?

2 K/U The balanced chemical equation for the formation of water from its elements is sometimes written as

$$H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(\ell)}$$

Explain why it is acceptable to use fractional coefficients in a balanced chemical equation.

Unit Investigation Prep

Before you design your quantitative analysis investigation at the end of Unit 2, decide how you will make use of the concepts you learned in this section. Assume that you know the identity of reactants and you know what products will be formed in the reaction. If you can measure how much product is formed in the reaction, can you determine how much reactant was initially present? Explain how, using an example.

3 C In the following reaction, does 1.0 g of sodium react completely with 0.50 g of chlorine? Explain your answer.

$$Na_{(s)} + \frac{1}{2}Cl_{2(g)} \rightarrow NaCl$$

4 K/U Sulfur and oxygen can combine to form sulfur dioxide, SO_2, and sulfur trioxide, SO_3.

(a) Write a balanced chemical equation for the formation of SO_2 from S and O_2.

(b) Write a balanced chemical equation for the formation of SO_3.

(c) How many moles of O_2 must react with 1 mol of S to form 1 mol of SO_3?

(d) What mass of O_2 is needed to react with 32.1 g of S to form SO_3?

5 K/U The balanced chemical equation for the combustion of propane is

$$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$$

(a) Write the mole ratios for the reactants and products in the combustion of propane.

(b) How many moles of O_2 are needed to react with 0.500 mol of C_3H_8?

(c) How many molecules of O_2 are needed to react with 2.00 mol of C_3H_8?

(d) If 3.00 mol of C_3H_8 burn completely in O_2, how many moles of CO_2 are produced?

6 I Phosphorus pentachloride, PCl_5, reacts with water to form phosphoric acid, H_3PO_4, and hydrochloric acid, HCl.

$$PCl_{5(s)} + 4H_2O_{(\ell)} \rightarrow H_3PO_{4(aq)} + 5HCl_{(aq)}$$

(a) What mass of PCl_5 is needed to react with an excess quantity of H_2O to produce 23.5 g of H_3PO_4?

(b) How many molecules of H_2O are needed to react with 3.87 g of PCl_5?

7 I A chemist has a beaker containing lead nitrate, $Pb(NO_3)_2$, dissolved in water. The chemist adds a solution containing sodium iodide, NaI, and a bright yellow precipitate is formed. The chemist continues to add NaI until no further yellow precipitate is formed. The chemist filters the precipitate, dries it in an oven, and finds it has a mass of 1.43 g.

(a) Write a balanced chemical equation to describe what happened in this experiment. Hint: compounds with sodium ions are always soluble.

(b) Use the balanced chemical equation to determine what mass of lead nitrate, $Pb(NO_3)_2$, was dissolved in the water in the beaker.

8 MC The Apollo-13 mission overcame an astonishing number of difficulties on its return to Earth. One problem the astronauts encountered was removing carbon dioxide from the air they were breathing. Do some research to find out:

(a) What happened to lead to an unexpected accumulation of carbon dioxide?

(b) What did the astronauts do to overcome this difficulty?

The Limiting Reactant

7.2

Section Preview/ Specific Expectations

In this section, you will

- **calculate,** for any given reactant or product in a chemical equation, the corresponding mass or quantity (in moles or molecules) of any other reactant or product
- **perform** an investigation to determine the limiting reactant in a chemical reaction
- **assess** the importance of determining the limiting reactant
- **solve** problems involving percentage yield and limiting reactants
- **communicate** your understanding of the following terms: *stoichiometric coefficients, stoichiometric amounts, limiting reactant, excess reactant*

A balanced chemical equation shows the mole ratios of the reactants and products. To emphasize this, the coefficients of equations are sometimes called **stoichiometric coefficients**. Reactants are said to be present in **stoichiometric amounts** when they are present in a mole ratio that corresponds exactly to the mole ratio predicted by the balanced chemical equation. This means that when a reaction is complete, there are no reactants left. In practice, however, there often *are* reactants left.

In the previous section, you looked at an "equation" for making a salad. You looked at situations in which you had the right amounts of ingredients to make one or more salads, with no leftover ingredients.

$$\text{1 head of lettuce} + \text{2 cucumbers} + \text{5 radishes} \rightarrow \text{1 salad}$$

What if you have two heads of lettuce, 12 cucumbers, and 25 radishes, as in Figure 7.5? How many salads can you make? Because each salad requires two heads of lettuce, you can make only two salads. Here the amount of lettuce limits the number of salads you can make. Some of the other two ingredients are left over.

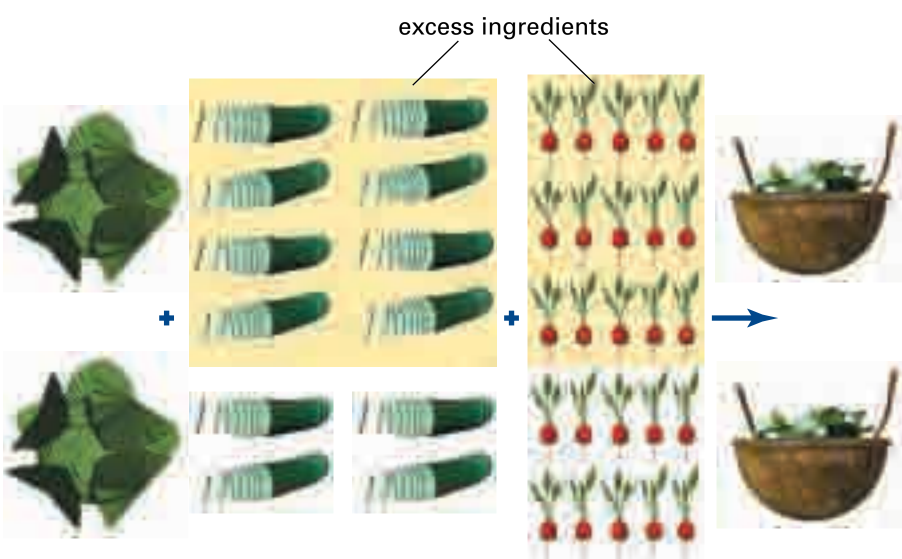


Figure 7.5 Which ingredient limits how many salads can be made?

Chemical reactions often work in the same way. For example, consider the first step in extracting zinc from zinc oxide:

$$ZnO_{(s)} + C_{(s)} \rightarrow Zn_{(s)} + CO_{(g)}$$

If you were carrying out this reaction in a laboratory, you could obtain samples of zinc oxide and carbon in a 1:1 mole ratio. In an industrial setting, however, it is impractical to spend time and money ensuring that zinc oxide and carbon are present in stoichiometric amounts. It is also unnecessary. In an industrial setting, engineers add more carbon, in the form of charcoal, than is necessary for the reaction. All the zinc oxide reacts, but there is carbon left over.

Figure 7.6 All the gasoline in this car's tank has reacted. Thus, even though there is still oxygen available in the air, the combustion reaction cannot proceed.

Having one or more reactants in excess is very common. Another example is seen in gasoline-powered vehicles. Their operation depends on the reaction between fuel and oxygen. Normally, the fuel-injection system regulates how much air enters the combustion chamber, and oxygen is the limiting reactant. When the fuel is very low, however, fuel becomes the limiting reactant and the reaction cannot proceed, as in Figure 7.6.

In nature, reactions almost never have reactants in stoichiometric amounts. Think about respiration, represented by the following chemical equation:

$$C_6H_{12}O_{6(s)} + 6O_{2(g)} \rightarrow 6CO_{2(g)} + 6H_2O_{(\ell)}$$

When an animal carries out respiration, there is an unlimited amount of oxygen in the air. The amount of glucose, however, depends on how much food the animal has eaten.

ThoughtLab The Limiting Item

Imagine that you are in the business of producing cars. A simplified "equation" for making a car is
1 car body + 4 wheels + 2 wiper blades → 1 car

Procedure

1. Assume that you have 35 car bodies, 120 wheels, and 150 wiper blades in your factory. How many complete cars can you make?
2. (a) Which item "limits" the number of complete cars that you can make? Stated another way, which item will "run out" first?
   (b) Which items are present in excess amounts?
   (c) How much of each "excess" item remains after the "reaction"?

Analysis

1. Does the *amount* that an item is in excess affect the quantity of the product that is made? Explain.
2. There are fewer car bodies than wheels and wiper blades. Explain why car bodies are not the limiting item, in spite of being present in the smallest amount.

Determining the Limiting Reactant

The reactant that is completely used up in a chemical reaction is called the **limiting reactant**. In other words, the limiting reactant determines how much product is produced. When the limiting reactant is used up, the reaction stops. In real-life situations, there is almost always a limiting reactant.

A reactant that remains after a reaction is over is called the **excess reactant**. Once the limiting reactant is used, no more product can be made, regardless of how much of the excess reactants may be present.

When you are given amounts of two or more reactants to solve a stoichiometric problem, you first need to identify the limiting reactant. One way to do this is to find out how much product would be produced by each reactant if the other reactant were present in excess. The reactant that produces the least amount of product is the limiting reactant. Examine the following Sample Problem to see how to use this approach to identify the limiting reactant.

Sample Problem

Identifying the Limiting Reactant

Problem

Lithium nitride reacts with water to form ammonia and lithium hydroxide, according to the following balanced chemical equation:

$$Li_3N_{(s)} + 3H_2O_{(\ell)} \rightarrow NH_{3(g)} + 3LiOH_{(aq)}$$

If 4.87 g of lithium nitride reacts with 5.80 g of water, find the limiting reactant.

What Is Required?

You need to determine whether lithium nitride or water is the limiting reactant.

What Is Given?

Reactant: lithium nitride, $Li_3N \rightarrow 4.87$ g
Reactant: water, $H_2O \rightarrow 5.80$ g
Product: ammonia, NH_3
Product: lithium hydroxide, LiOH

Plan Your Strategy

Convert the given masses into moles. Use the mole ratios of reactants and products to determine how much ammonia is produced by each amount of reactant. The limiting reactant is the reactant that produces the smaller amount of product.

Act on Your Strategy

$$n \text{ mol } Li_3N = \frac{4.87 \cancel{\text{ g}} \ Li_3N}{34.8 \cancel{\text{ g}}/\text{mol}}$$
$$= 0.140 \text{ mol } Li_3N$$

$$n \text{ mol } H_2O = \frac{5.80 \cancel{\text{ g}} \ H_2O}{18.0 \cancel{\text{ g}}/\text{mol}}$$
$$= 0.322 \text{ mol } H_2O$$

Calculate the amount of NH_3 produced, based on the amount of Li_3N.

$$n \text{ mol of } NH_3 = \frac{1 \text{ mol } NH_3}{1 \cancel{\text{ mol } Li_3N}} (0.140 \cancel{\text{ mol } Li_3N})$$
$$= 0.140 \text{ mol } NH_3$$

Calculate the amount of NH_3 produced, based on the amount of H_2O.

Continued ...

PROBLEM TIP

To determine the limiting reactant, you can calculate how much of either ammonia or lithium hydroxide would be produced by the reactants. In this problem, ammonia was chosen because only one mole is produced, simplifying the calculation.

Continued ...
FROM PAGE 253

$$n \text{ mol } NH_3 = \frac{1 \text{ mol } NH_3}{3 \cancel{\text{ mol } H_2O}} \times (0.322 \cancel{\text{ mol } H_2O})$$
$$= 0.107 \text{ mol } NH_3$$

The water would produce less ammonia than the lithium nitride. Therefore, the limiting reactant is water. Notice that there is more water than lithium nitride, in terms of mass and moles. Water is the limiting reactant, however, because 3 mol of water are needed to react with 1 mol of lithium nitride.

Check Your Solution

According to the balanced chemical equation, the ratio of lithium nitride to water is 1/3. The ratio of lithium nitride to water, based on the mole amounts calculated, is 0.14:0.32. Divide this ratio by 0.14 to get 1.0:2.3. For each mole of lithium nitride, there are only 2.3 mol water. However, 3 mol are required by stoichiometry. Therefore, water is the limiting reactant.

Practice Problems

23. The following balanced chemical equation shows the reaction of aluminum with copper(II) chloride. If 0.25 g of aluminum reacts with 0.51 g of copper(II) chloride, determine the limiting reactant.

$$2Al_{(s)} + 3CuCl_{2(aq)} \rightarrow 3Cu_{(s)} + 2AlCl_{3(aq)}$$

$\frac{2}{3}$ | $\frac{3}{4.5}$ | $\frac{3}{1}$

AL $\frac{0.25}{27} = 0.00925$

$CuCl_2 = \frac{0.519}{63.54+35.45 \times 2}$

$= \frac{0.51}{134.446} =$

$= 0.0038$.

0.00925 AL 3

0.0038 $CuCl_2$ 1

AL 的消耗:

$\frac{0.0038}{3} \times 2 = 0.0025$

AL 的 remain:

$0.00925 - 0.0025 = 0.00675$

24. Hydrogen fluoride, HF, is a highly toxic gas. It is produced by the double displacement reaction of calcium fluoride, CaF_2, with concentrated sulfuric acid, H_2SO_4.

$$CaF_{2(s)} + H_2SO_{4(\ell)} \rightarrow 2HF_{(g)} + CaSO_{4(s)}$$

Determine the limiting reactant when 10.0 g of CaF_2 reacts with 15.5 g of H_2SO_4.

25. Acrylic, a common synthetic fibre, is formed from acrylonitrile, C_3H_3N. Acrylonitrile can be prepared by the reaction of propylene, C_3H_6, with nitric oxide, NO.

$$4C_3H_{6(g)} + 6NO_{(g)} \rightarrow 4C_3H_3N_{(g)} + 6H_2O_{(g)} + N_{2(g)}$$

What is the limiting reactant when 126 g of C_3H_6 reacts with 175 g of NO?

26. 3.76 g of zinc reacts with 8.93×10^{23} molecules of hydrogen chloride. Which reactant is present in excess?

You now know how to use a balanced chemical equation to find the limiting reactant. Can you find the limiting reactant by experimenting? You know that the limiting reactant is completely consumed in a reaction, while any reactants in excess remain after the reaction is finished. In Investigation 7-A, you will observe a reaction and identify the limiting reactant, based on your observations.

Investigation 7-A

SKILL FOCUS

- Predicting
- Performing and recording
- Analyzing and interpreting
- Communicating results

Limiting and Excess Reactants

In this investigation, you will predict and observe a limiting reactant. You will use the single replacement reaction of aluminum with aqueous copper(II) chloride:

$$2Al_{(s)} + 3CuCl_{2(aq)} \rightarrow 3Cu_{(s)} + 2AlCl_{3(aq)}$$

Note that copper(II) chloride, $CuCl_2$, is light blue in aqueous solution. This is due to the $Cu^{2+}{}_{(aq)}$ ion. Aluminum chloride, $AlCl_{3(aq)}$, is colourless in aqueous solution.

Question

How can observations tell you which is the limiting reactant in the reaction of aluminum with aqueous copper(II) chloride?

Prediction

Your teacher will give you a beaker that contains a 0.25 g piece of aluminum foil and 0.51 g of copper(II) chloride. Predict which one of these reactants is the limiting reactant.

Materials

100 mL beaker or 125 mL Erlenmeyer flask
stirring rod
0.51 g $CuCl_2$
0.25 g Al foil

Safety Precautions

The reaction mixture may get hot. Do not hold the beaker as the reaction proceeds.

Procedure

1. To begin the reaction, add about 50 mL of water to the beaker that contains the aluminum foil and copper(II) chloride.
2. Record the colour of the solution and any metal that is present at the beginning of the reaction.
3. Record any colour changes as the reaction proceeds. Stir occasionally with the stirring rod.
4. When the reaction is complete, return the beaker, with its contents, to your teacher for proper disposal. Do not pour anything down the drain.

Analysis

1. According to your observations, which reactant was present in excess? Which reactant was the limiting reactant?
2. How does your prediction compare with your observations?
3. Do stoichiometric calculations to support your observations of the limiting reactant. Refer to the previous ThoughtLab if you need help.
4. If your prediction of the limiting reactant was incorrect, explain why.

Conclusions

5. Write a conclusion to explain how your experimental observation supported your theoretical calculations.

Applications

6. Magnesium ($Mg_{(s)}$) and hydrogen chloride ($HCl_{(aq)}$) react according to the following skeleton equation:

$$Mg_{(s)} + HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

   (a) Balance the skeleton equation.
   (b) Examine the equation carefully. What evidence would you have that a reaction was taking place between the hydrochloric acid and the magnesium?
   (c) You have a piece of magnesium of unknown mass, and a beaker of water in which is dissolved an unknown amount of hydrogen chloride. Design an experiment to determine which reactant is the limiting reactant.

Write a balanced chemical equation.

↓

Identify the limiting reactant. Express it as an amount in moles.

↓

Calculate the amount of the required substance based on the amount of the limiting reactant.

↓

Convert the amount of the required substance to mass or number of particles, as directed by the question.

Figure 7.7 **Be sure to determine the limiting reactant in any stoichiometric problem before you solve it.**

PROBEWARE

If you have access to probeware, do the Chemistry 11 lab, Stoichiometry, now.

The Limiting Reactant in Stoichiometric Problems

You are now ready to use what you know about finding the limiting reactant to predict the amount of product that is expected in a reaction. This type of prediction is a routine part of a chemist's job, both in academic research and industry. To produce a compound, for example, chemists need to know how much product they can expect from a given reaction. In analytical chemistry, chemists often analyze an impure substance by allowing it to react in a known reaction. They predict the expected mass of the product(s) and compare it with the actual mass of the product(s) obtained. Then they can determine the purity of the compound.

Since chemical reactions usually occur with one or more of the reactants in excess, you often need to determine the limiting reactant before you carry out stoichiometric calculations. You can incorporate this step into the process you have been using to solve stoichiometric problems, as shown in Figure 7.7.

Sample Problem

The Limiting Reactant in a Stoichiometric Problem

Problem

White phosphorus consists of a molecule made up of four phosphorus atoms. It burns in pure oxygen to produce tetraphosphorus decaoxide.

$$P_{4(s)} + 5O_{2(g)} \rightarrow P_4O_{10(s)}$$

A 1.00 g piece of phosphorus is burned in a flask filled with 2.60×10^{23} molecules of oxygen gas. What mass of tetraphosphorus decaoxide is produced?

What Is Required?

You need to find the mass of tetraphosphorus decaoxide that is produced.

What Is Given?

You know the balanced chemical equation. You also know the mass of phosphorus and the number of oxygen molecules that reacted.

Plan Your Strategy

First convert each reactant to moles and find the limiting reactant. Using the mole to mole ratio of the limiting reactant to the product, determine the number of moles of tetraphosphorus decaoxide that is expected. Convert this number of moles to grams.

Act on Your Strategy

$$\begin{aligned} n \text{ mol } P_4 &= \frac{1.00 \cancel{\text{g}}\ P_4}{123.9 \cancel{\text{g}}/\text{mol } P_4} \\ &= 8.07 \times 10^{-3} \text{ mol } P_4 \end{aligned}$$

Continued ...

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$$n \text{ mol } O_2 = \frac{2.60 \times 10^{23} \cancel{\text{ molecules}}}{6.02 \times 10^{23} \cancel{\text{ molecules}}/\text{mol}}$$
$$= 0.432 \text{ mol } O_2$$

Calculate the amount of P_4O_{10} that would be produced by the P_4.

$$\frac{n \text{ mol } P_4O_{10}}{8.07 \times 10^{-3} \text{ mol } P_4} = \frac{1 \text{ mol } P_4O_{10}}{1 \text{ mol } P_4}$$
$$(\cancel{8.07 \times 10^{-3} \text{ mol } P_4})\frac{n \text{ mol } P_4O_{10}}{\cancel{8.07 \times 10^{-3} \text{ mol } P_4}} = \frac{1 \text{ mol } P_4O_{10}}{1 \cancel{\text{ mol } P_4}}(8.07 \times 10^{-3} \cancel{\text{ mol } P_4})$$
$$= 8.07 \times 10^{-3} \text{ mol } P_4O_{10}$$

Calculate the amount of P_4O_{10} that would be produced by the O_2.

$$\frac{n \text{ mol } P_4O_{10}}{0.432 \text{ mol } O_2} = \frac{1 \text{ mol } P_4O_{10}}{5 \text{ mol } O_2}$$
$$(\cancel{0.432 \text{ mol } O_2})\frac{n \text{ mol } P_4O_{10}}{\cancel{0.432 \text{ mol } O_2}} = \frac{1 \text{ mol } P_4O_{10}}{5 \cancel{\text{ mol } O_2}}(0.432 \cancel{\text{ mol } O_2})$$
$$= 8.64 \times 10^{-2} \text{ mol } P_4O_{10}$$

Since P_4 would produce less P_4O_{10} than O_2 would, P_4 is the limiting reactant.

$$P_{4(s)} + 5O_{2(g)} \longrightarrow P_4O_{10(s)}$$

0.00807 mol → 0.00807 mol

unknown ratio **known ratio**

$$\frac{n \text{ mol } P_4O_{10}}{0.00807 \text{ mol } P_4} = \frac{1 \text{ mol } P_4O_{10}}{1 \text{ mol } P_4}$$
$$(0.00807 \cancel{\text{ mol } P_4})\frac{n \text{ mol } P_4O_{10}}{\cancel{0.00807 \text{ mol } P_4}} = (0.00807 \cancel{\text{ mol } P_4})\frac{1 \text{ mol } P_4O_{10}}{1 \cancel{\text{ mol } P_4}}$$
$$= 0.00807 \text{ mol } P_4O_{10}$$

0.00807 mol P_4O_{10}
× 284 g/mol P_4O_{10}

↓

2.29 g P_4O_{10}

Check Your Solution

There were more than 5 times as many moles of O_2 as moles of P_4, so it makes sense that P_4 was the limiting reactant. An expected mass of 2.29 g of tetraphosphorus decaoxide is reasonable. It is formed in a 1:1 ratio from phosphorus. It has a molar mass that is just over twice the molar mass of phosphorus.

Practice Problems

27. Chloride dioxide, ClO_2, is a reactive oxidizing agent. It is used to purify water.

$$6ClO_{2(g)} + 3H_2O_{(\ell)} \rightarrow 5HClO_{3(aq)} + HCl_{(aq)}$$

(a) If 71.00 g of ClO_2 is mixed with 19.00 g of water, what is the limiting reactant?

(b) What mass of $HClO_3$ is expected in part (a)?

(c) How many molecules of HCl are expected in part (a)?

Continued ...

Continued ...
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28. Hydrazine, N_2H_4, reacts exothermically with hydrogen peroxide, H_2O_2

$$N_2H_{4(\ell)} + 7H_2O_{2(aq)} \rightarrow 2HNO_{3(g)} + 8H_2O_{(g)}$$

(a) 120 g of N_2H_4 reacts with an equal mass of H_2O_2. Which is the limiting reactant?

(b) What mass of HNO_3 is expected?

(c) What mass, in grams, of the excess reactant remains at the end of the reaction?

29. In the textile industry, chlorine is used to bleach fabrics. Any of the toxic chlorine that remains after the bleaching process is destroyed by reacting it with a sodium thiosulfate solution, $Na_2S_2O_{3(aq)}$.

$$Na_2S_2O_{3(aq)} + 4Cl_{2(g)} + 5H_2O_{(\ell)} \rightarrow 2NaHSO_{4(aq)} + 8HCl_{(aq)}$$

135 kg of $Na_2S_2O_3$ reacts with 50.0 kg of Cl_2 and 238 kg of water. How many grams of $NaHSO_4$ are expected?

30. Manganese(III) fluoride can be formed by the reaction of manganese(II) iodide with fluorine.

$$2MnI_{2(s)} + 13F_{2(g)} \rightarrow 2MnF_{3(s)} + 4IF_{5(\ell)}$$

(a) 1.23 g of MnI_2 reacts with 25.0 g of F_2. What mass of MnF_3 is expected?

(b) How many molecules of IF_5 are produced in part (a)?

(c) What reactant is in excess? How much of it remains at the end of the reaction?

COURSE CHALLENGE

You will use the concepts of stoichiometry and limiting reactants in the Chemistry Course Challenge. If you have two reactants and you want to use up all of one reactant, which is the limiting reactant?

CHEM FACT

Carbon disulfide, CS_2, is an extremely volatile and flammable substance. It is so flammable that it can ignite when exposed to boiling water! Because carbon disulfide vapour is more than twice as dense as air, it can "blanket" the floor of a laboratory. There have been cases where the spark from an electrical motor has ignited carbon disulfide vapour in a laboratory, causing considerable damage. For this reason, specially insulated electrical motors are required in laboratory refrigerators and equipment.

Section Wrap-up

You now know how to identify a limiting reactant. This allows you to predict the amount of product that will be formed in a reaction. Often, however, your prediction will not accurately reflect reality. When a chemical reactions occurs—whether in a laboratory, in nature, or in industry—the amount of product that is formed is often different from the amount that was predicted by stoichiometric calculations. You will learn why this happens, and how chemists deal with it, in section 7.3.

Section Review

1. C Why do you not need to consider reactants that are present in excess amounts when carrying out stoichiometric calculations? Use an everyday analogy to explain the idea of excess quantity.

2. **(a)** C Magnesium reacts with oxygen gas, O_2, from the air. Which reactant do you think will be present in excess?

(b) C Gold is an extremely unreactive metal. Gold does react, however, with *aqua regia* (a mixture of concentrated nitric acid, $HNO_{3(aq)}$, and hydrochloric acid, $HCl_{(aq)}$). The complex ion $AuCl_4^-$, as well as NO_2 and H_2O, are formed. This reaction is always carried out with *aqua regia* in excess. Why would a chemist not have the gold in excess?

(c) C In general, what characteristics or properties of a chemical compound or atom make it suitable to be used as an *excess* reactant?

3 I Copper is a relatively inert metal. It is unreactive with most acids. It does, however, react with nitric acid.

$$3Cu_{(s)} + 8HNO_{3(aq)} \rightarrow 3Cu(NO_3)_{2(aq)} + 2NO_{(g)} + 4H_2O_{(\ell)}$$

What mass of NO is produced when 57.4 g of Cu reacts with 165 g of HNO_3?

4 I Iron can be produced when iron(III) oxide reacts with carbon monoxide gas.

$$Fe_2O_{3(s)} + 3CO_{(g)} \rightarrow 2Fe_{(s)} + 3CO_{2(g)}$$

11.5 g of Fe_2O_3 reacts with 2.63×10^{24} molecules of CO. What mass of Fe is expected?

5 I The reaction of an aqueous solution of iron(III) sulfate with aqueous sodium hydroxide produces aqueous sodium sulfate and a solid precipitate, iron(III) hydroxide.

$$Fe_2(SO_4)_{3(aq)} + 6NaOH_{(aq)} \rightarrow 3Na_2SO_{4(aq)} + 2Fe(OH)_{3(s)}$$

What mass of $Fe(OH)_3$ is produced when 10.0 g of $Fe_2(SO_4)_3$ reacts with an equal mass of NaOH?

6 I Carbon disulfide is used as a solvent for water-insoluble compounds, such as fats, oils, and waxes. Calculate the mass of carbon disulfide that is produced when 17.5 g of carbon reacts with 225 g of sulfur dioxide according to the following equation:

$$5C_{(s)} + 2SO_{2(g)} \rightarrow CS_{2(\ell)} + 4CO_{(g)}$$

7 I A chemist adds some zinc shavings to a beaker containing a blue solution of copper chloride. The contents of the beaker are stirred. After about an hour, the chemist observes that the blue colour has not completely disappeared.

(a) Write a balanced chemical equation to describe this reaction.

(b) What other observations would you expect the chemist to make?

(c) According to the chemist's observations, which reactant was the limiting reactant?

(d) The beaker contained 3.12 g of copper chloride dissolved in water. What does this tell you, quantitatively, about the amount of zinc that was added?

Nitric acid reacts with copper metal to produce poisonous, brown nitrogen dioxide, NO_2, gas.

Unit Investigation Prep

Consider what you have learned about limiting reactants when you design your quantitative analysis experiment at the end of Unit 2. Imagine you add one reactant (A) to an unknown amount of a second reactant (B). You intend to analyze the products (C) in order to calculate the amount of B. In this case, which reactant should be the limiting reactant, A or B? How do you know which reactant is the limiting reactant when you do not know the amount of reactant B?

7.3 Percentage Yield

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Section Preview/ Specific Expectations

In this section, you will

- **solve** problems involving percentage yield and limiting reactants
- **compare**, using laboratory results, the theoretical yield of a reaction with the actual yield
- **calculate** the percentage yield of a reaction, and suggest sources of experimental error
- **solve** stoichiometric problems involving the percentage purity of the reactants
- **communicate** your understanding of the following terms: *theoretical yield, actual yield, competing reaction, percentage yield, percentage purity*

When you write an examination, the highest grade that you can earn is usually 100%. Most people, however, do not regularly earn a grade of 100%. A percentage on an examination is calculated using the following equation:

$$\text{Percentage grade} = \frac{\text{Marks earned}}{\text{Maximum possible marks}} \times 100\%$$

Similarly, a batter does not succeed at every swing. A batter's success rate is expressed as a decimal fraction. The decimal can be converted to a percent by multiplying by 100%, as shown in Figure 7.8. In this section, you will learn about a percentage that chemists use to predict and express the "success" of reactions.

Figure 7.8 A baseball player's batting average is calculated as hits/attempts. For example, a player with 6 hits for 21 times at bat has a batting average of 6/21 = 0.286. This represents a success rate of 28.6%.

Theoretical Yield and Actual Yield

Chemists use stoichiometry to predict the amount of product that can be expected from a chemical reaction. The amount of product that is predicted by stoichiometry is called the **theoretical yield**. This predicted yield, however, is not always the same as the amount of product that is actually obtained from a chemical reaction. The amount of product that is obtained in an experiment is called the **actual yield**.

CHEM FACT

Actual yield is a *measured* quantity. Theoretical yield is a *calculated* quantity.

Why Actual Yield and Theoretical Yield Are Often Different

The actual yield of chemical reactions is usually less than the theoretical yield. This is caused by a variety of factors. For example, sometimes less than perfect collection techniques contribute to a lower than expected yield.

A reduced yield may also be caused by a **competing reaction**: a reaction that occurs at the same time as the principal reaction and involves its reactants and/or products. For example, phosphorus reacts with chlorine to form phosphorus trichloride. Some of the phosphorus trichloride, however, can then react with chlorine to form phosphorus pentachloride.

Here are the chemical equations for these competing reactions:

$$2P_{(s)} + 3Cl_{2(g)} \rightarrow 2PCl_{3(\ell)}$$
$$PCl_{3(\ell)} + Cl_{2(g)} \rightarrow PCl_{5(s)}$$

Therefore, not all the phosphorus is converted to phosphorus trichloride. So the actual yield of phosphorus trichloride is less than the theoretical yield.

Experimental design and technique may affect the actual yield, as well. For example, suppose that you need to obtain a product by filtration. Some of the product may remain in solution and therefore not be caught on the filter paper.

Another common cause of reduced yield is impure reactants. The theoretical yield is calculated based on the assumption that reactants are pure. You will learn about the effects of impure reactants on page 265.

Calculating Percentage Yield

The **percentage yield** of a chemical reaction compares the mass of product obtained by experiment (the actual yield) with the mass of product determined by stoichiometric calculations (the theoretical yield). It is calculated as follows:

$$\text{Percentage yield} = \left(\frac{\text{Actual yield}}{\text{Theoretical yield}}\right) \times 100\%$$

In section 7.1, you looked at the reaction of hydrogen and nitrogen to produce ammonia. You assumed that all the nitrogen and hydrogen reacted. Under certain conditions of temperature and pressure, this is a reasonable assumption. When ammonia is produced industrially, however, temperature and pressure are manipulated to maximize the speed of production. Under these conditions, the actual yield is much less than the theoretical yield. Examine the next Sample Problem to learn how to calculate percentage yield.

Simple

Sample Problem

Calculating Percentage Yield

Problem

氨气.

Ammonia can be prepared by reacting nitrogen gas, taken from the atmosphere, with hydrogen gas.

5.2 mol × (14+3)

2.61 mol 75g $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$ 88.4g

When 7.5×10^1 g of nitrogen reacts with sufficient hydrogen, the theoretical yield of ammonia is 9.10 g. (You can verify this by doing the stoichiometric calculations.) If 1.72 g of ammonia is obtained by experiment, what is the percentage yield of the reaction?

$\frac{9.1}{88.4} \times 100\% = 10.3\%$

What Is Required?

You need to find the percentage yield of the reaction.

What Is Given?

:N≡N:

actual yield = 1.72 g
theoretical yield = 9.10 g

Continued ...

Continued ...

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Plan Your Strategy

Divide the actual yield by the theoretical yield, and multiply by 100%.

Act on Your Strategy

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$
$$= \frac{1.72\ \cancel{g}}{9.10\ \cancel{g}} \times 100\%$$
$$= 18.9\%$$

The percentage yield of the reaction is 18.9%.

Check Your Solution

By inspection, you can see that 1.72 g is roughly 20% of 9.10 g.

Practice Problems

31. 20.0 g of bromic acid, $HBrO_3$, is reacted with excess HBr.

$$HBrO_{3(aq)} + 5HBr_{(aq)} \rightarrow 3H_2O_{(\ell)} + 3Br_{2(aq)}$$

(a) What is the theoretical yield of Br_2 for this reaction?

(b) If 47.3 g of Br_2 is produced, what is the percentage yield of Br_2?

32. Barium sulfate forms as a precipitate in the following reaction:

$$Ba(NO_3)_{2(aq)} + Na_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + 2NaNO_{3(aq)}$$

When 35.0 g of $Ba(NO_3)_2$ is reacted with excess Na_2SO_4, 29.8 g of $BaSO_4$ is recovered by the chemist.

(a) Calculate the theoretical yield of $BaSO_4$.

(b) Calculate the percentage yield of $BaSO_4$.

33. Yeasts can act on a sugar, such as glucose, $C_6H_{12}O_6$, to produce ethyl alcohol, C_2H_5OH, and carbon dioxide.

$$C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$$

If 223 g of ethyl alcohol are recovered after 1.63 kg of glucose react, what is the percentage yield of the reaction?

Sometimes chemists know what percentage yield to expect from a chemical reaction. This is especially true of an industrial reaction, where a lot of experimental data are available. As well, the reaction has usually been carried out many times, with large amounts of reactants. Examine the next Sample Problem to learn how to predict the actual yield of a reaction from a known percentage yield.

Sample Problem

Predicting Actual Yield Based on Percentage Yield

Problem

Calcium carbonate can be thermally decomposed to calcium oxide and carbon dioxide.

$$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$$

Under certain conditions, this reaction proceeds with a 92.4% yield of calcium oxide. How many grams of calcium oxide can the chemist expect to obtain if 12.4 g of calcium carbonate is heated?

What Is Required?

You need to calculate the amount of calcium oxide, in grams, that will be formed in the reaction.

What Is Given?

Percentage yield CaO = 92.4%
m $CaCO_3$ = 12.4 g

Plan Your Strategy

Calculate the theoretical yield of calcium oxide using stoichiometry. Then multiply the theoretical yield by the percentage yield to predict the actual yield.

Act on Your Strategy

1 $$CaCO_{3(s)} \longrightarrow CaO_{(s)} + CO_{2(g)}$$

0.124 mol → 0.124 mol

3 **unknown ratio** **known ratio**

$$\frac{\text{amount CaCO}}{0.124 \text{ mol } CaCO_3} = \frac{1 \text{ mol CaO}}{1 \text{ mol } CaO_3}$$

$$(\cancel{0.124 \text{ mol } CaCO_3})\frac{\text{amount CaO}}{\cancel{0.124 \text{ mol } CaCO_3}} = (0.124 \text{ mol } \cancel{CaCO_3})\frac{1 \text{ mol CaO}}{1 \text{ mol } \cancel{CaO_3}}$$

$$= 0.124 \text{ mol CaO}$$

2 $$\frac{12.4 \text{ g } \cancel{CaCO_3}}{100 \text{ g } \cancel{CaCO_3}/\text{mol } CaCO_3} = 0.124 \text{ mol } CaCO_3$$

4 $$0.124 \cancel{\text{ mol CaO}} \times 56.1 \text{ g CaO}/\cancel{\text{mol CaO}} = 6.95 \text{ g CaO}$$

5 $$\text{Actual yield} = 6.95 \text{ g CaO} \times \frac{92.4}{100} = 6.42 \text{ g CaO}$$

12.4 g $CaCO_3$ 6.95 g CaO

Check Your Solution

92.5% of 6.95 g is about 6.4 g. The answer is reasonable.

Continued ...

Continued ...
FROM PAGE 263

Practice Problems

34. The following reaction proceeds with a 70% yield.

$$C_6H_{6(\ell)} + HNO_{3(aq)} \rightarrow C_6H_5NO_{2(\ell)} + H_2O_{(\ell)}$$

Calculate the mass of $C_6H_5NO_2$ expected if 12.8 g of C_6H_6 reacts with excess HNO_3.

35. The reaction of toluene, C_7H_8, with potassium permanganate, $KMnO_4$, gives less than a 100% yield.

$$C_7H_{8(\ell)} + 2KMnO_{4(aq)} \rightarrow KC_7H_5O_{2(aq)} + 2MnO_{2(s)} + KOH_{(aq)} + H_2O_{(\ell)}$$

(a) 8.60 g of C_7H_8 is reacted with excess $KMnO_4$. What is the theoretical yield, in grams, of $KC_7H_5O_2$?

(b) If the percentage yield is 70.0%, what mass of $KC_7H_5O_2$ can be expected?

(c) What mass of C_7H_8 is needed to produce 13.4 g of $KC_7H_5O_2$, assuming a yield of 60%?

36. Marble is made primarily of calcium carbonate. When calcium carbonate reacts with hydrogen chloride, it reacts to form calcium chloride, carbon dioxide and water. If this reaction occurs with 81.5% yield, what mass of carbon dioxide will be collected if 15.7 g of $CaCO_3$ is added to sufficient hydrogen chloride?

37. Mercury, in its elemental form or in a chemical compound is highly toxic. Water-soluble mercury compounds, such as mercury(II) nitrate, can be removed from industrial wastewater by adding sodium sulfide to the water, which forms a precipitate of mercury(II) sulfide, which can then be filtered out.

$$Hg(NO_3)_{2(aq)} + Na_2S_{(aq)} \rightarrow HgS_{(s)} + 2NaNO_{3(aq)}$$

If 3.45×10^{23} formula units of $Hg(NO_3)_2$ are reacted with excess Na_2S, what mass of HgS can be expected if this process occurs with 97.0% yield?

mind STRETCH

You hear a great deal about the fuel consumption of automobiles. What about air consumption? Your challenge is to determine the information you need to answer the following question: What mass of air does an automobile require to travel from Thunder Bay, Ontario, to Smooth Rock Falls, Ontario? This is a distance of 670 km.

When you are finished, go to question 23 on page 273. You can check your answer and solve the problem, too.

COURSE CHALLENGE

How would you determine the percentage yield of a double displacement reaction that produces a precipitate? Consider this question to prepare for your Chemistry Course Challenge.

Applications of Percentage Yield

The percentage yield of chemical reactions is extremely important in industrial chemistry and the pharmaceutical industry. For example, the synthesis of certain drugs involves many sequential chemical reactions. Often each reaction has a low percentage yield. This results in a tiny overall yield. Research chemists, who generally work with small quantities of reactants, may be satisfied with a poor yield. Chemical engineers, on the other hand, work with very large quantities. They may use hundreds or even thousands of kilograms of reactants! A difference of 1% in the yield of a reaction can translate into thousands of dollars.

The work of a chemist in a laboratory can be likened to making spaghetti for a family. The work of a chemical engineer, by contrast, is like making spaghetti for 10 000 people! Learn more about chemical engineers in Careers in Chemistry on the next page. Then perform an investigation to determine the percentage yield of a reaction on page 266.

Careers in Chemistry

Chemical Engineer

Chemical engineers are sometimes described as "universal engineers" because of their unique knowledge of math, physics, engineering, and chemistry. This broad knowledge allows them to work in a variety of areas, from designing paint factories to developing better tasting, more nutritious foods. Canadian chemical engineers are helping to lead the world in making cheap, long-lasting, and high-quality CDs and DVDs. In addition to designing and operating commercial plants, chemical engineers can be found in university labs, government agencies, and consulting firms.

Producing More for Less

Once chemists have developed a product in a laboratory, it is up to chemical engineers to design a process to make the product in commercial quantities as efficiently as possible. "Scaling up" production is not just a matter of using larger beakers. Chemical engineers break down the chemical process into a series of smaller "unit operations" or processes and techniques. They use physics, chemistry, and complex mathematical models. For example, making liquid pharmaceutical products (such as syrups, solutions, and suspensions) on a large scale involves adding specific amounts of raw materials to large mixing tanks. Then the raw materials are heated to a set temperature and mixed at a set speed for a given amount of time. The final product is filtered and stored in holding tanks. Chemical engineers ensure that each process produces the maximum amount of product.

Becoming a Chemical Engineer

To become a chemical engineer, you need a bachelor's degree in chemical engineering. Most provinces also require a Professional Engineer (P. Eng.) designation. Professional engineers must have at least four years of experience and must pass an examination. As well, they must commit to continuing their education to keep up with current developments. Chemical engineers must be able to work well with people and to communicate well.

Make Career Connections

1. Discuss engineering studies and careers with working engineers, professors, and engineering students. Look for summer internship programs and job shadowing opportunities. Browse the Internet. Contact your provincial engineering association, engineering societies, and universities for more information.
2. Participate in National Engineering Week in Canada in March of each year. This is when postsecondary institutions, companies, science centres, and other organizations hold special events, including engineering contests and workshops.

Percentage Purity

Often impure reactants are the cause of a percentage yield of less than 100%. Impurities cause the mass data to be incorrect. For example, suppose that you have 1.00 g of sodium chloride and you want to carry out a reaction with it. You think that the sodium chloride may have absorbed some water, so you do not know exactly how much pure sodium chloride you have. If you calculate a theoretical yield for your reaction based on 1.00 g of sodium chloride, your actual yield will be less. There is not 1.00 g of sodium chloride in the sample.

Investigation 7-B

SKILL FOCUS

- Predicting
- Performing and recording
- Analyzing and interpreting

Determining the Percentage Yield of a Chemical Reaction

The percentage yield of a reaction is determined by numerous factors: The nature of the reaction itself, the conditions under which the reaction was carried out, and the nature of the reactants used.

In this investigation, you will determine the percentage yield of the following chemical reaction:

$$Fe_{(s)} + CuCl_{2(aq)} \rightarrow FeCl_{2(aq)} + Cu_{(s)}$$

You will use steel wool, since it is virtually pure iron.

Question

What is the percentage yield of the reaction of iron and copper chloride when steel wool and copper chloride dihydrate are used as reactants?

Predictions

Predict the mass of copper that will be produced if 1.00 g of iron (steel wool) reacts *completely* with a solution containing excess $CuCl_2$. Also predict the maximum possible yield.

Materials

2 beakers (250 mL)
stirring rod
electronic balance, accurate to two decimal places
distilled water
wash bottle with distilled water
drying oven or heat lamp
about 1.00 g rust-free, degreased steel wool
5.00 g copper chloride dihydrate, $CuCl_2{\cdot}2H_2O$
15 mL 1 mol/L hydrochloric acid, HCl

Safety Precautions

If you get either $CuCl_2$ or the HCl solution on your skin, flush with plenty of cold water.

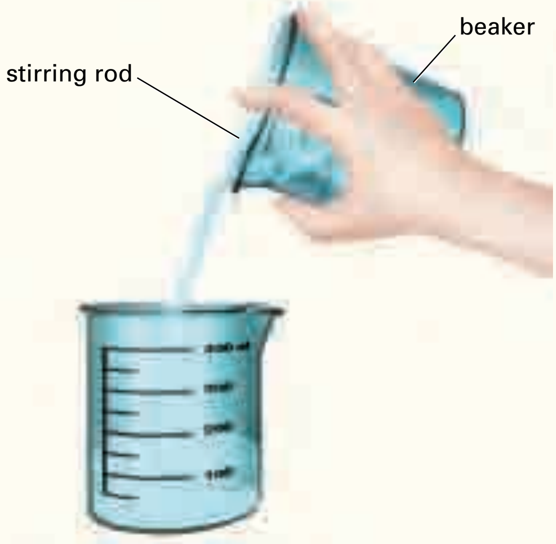
Procedure

1. Label a clean, dry 250 mL beaker with your initials. Use a glass marker, or write with pencil on the frosted area of the beaker. Do not use tape, since the beaker will be dried in an oven later.
2. Copy the table below into your notebook. Record the mass of the labelled beaker in your table.

Observations

Mass of empty beaker	
Mass of steel wool	
Mass of beaker containing clean, dry copper	

3. Put about 50 mL of distilled water in the beaker. Add 5.00 g of $CuCl_2{\cdot}2H_2O$ to the water. Stir to dissolve.
4. Record the mass of the steel wool in your table.
5. Add the steel wool to the $CuCl_2$ solution in the beaker. Allow it to sit until all the steel wool has reacted. This could take up to 20 min.
6. When the reaction is complete, decant the solution into a 250 mL beaker, as shown in the diagram.



Pouring *down* a stirring rod ensures that no liquid dribbles down the outside of the beaker. The glove in this illustration is omitted so you can clearly see where to place your fingers. Always wear gloves when handling chemicals in the laboratory.

7. Using a wash bottle, rinse the copper several times with distilled water. Decant the water as shown in the diagram.
8. Add 10 to 15 mL of 1 mol/L HCl to further wash the copper. Decant the HCl, and wash the copper again with distilled water. (If the copper is still not clean, wash it again with the HCl. Remember to do a final wash with distilled water.)
9. Place your labelled reaction beaker, containing the cleaned copper, in a drying oven overnight.
10. Find the mass of the beaker containing the dry copper.
11. Return the beaker, containing the copper, to your teacher for proper disposal.

Analysis

1. **(a)** Using the mass of the iron (steel wool) you used, calculate the theoretical yield of the copper, in grams.
   (b) How does the mass of the copper you collected compare with the expected theoretical yield?
2. Based on the amount of iron that you used, prove that the 5.00 g of $CuCl_2 \cdot 2H_2O$ was the excess reactant.

Conclusion

3. Calculate the percentage yield for this reaction.

Applications

4. If your percentage yield was not 100%, suggest sources of error.
5. How would you attain an improved percentage yield if you performed this reaction again? Consider your technique and materials.
6. Do some research to find out the percent by mass of iron in steel wool. Predict what your percentage yield would be if you had used pure iron in this reaction. Would it make a difference?

Figure 7.9 Copper is removed from mines like this one in the form of an ore. There must be sufficient copper in the ore to make the mine economically viable.

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In the mining industry, metals are usually recovered in the form of an ore. An ore is a naturally occurring rock that contains a high concentration of one or more metals. Whether an ore can be profitably mined depends on several factors: the cost of mining and refining the ore, the price of the extracted metal, and the cost of any legal and environmental issues related to land use. The inaccurate chemical analysis of an ore sample can cost investors millions of dollars if the ore deposit does not yield what was expected.

The **percentage purity** of a sample describes what proportion, by mass, of the sample is composed of a specific compound or element. For example, suppose that a sample of gold has a percentage purity of 98%. This means that every 100 g of the sample contains 98 g of gold and 2 g of impurities.

You can apply your knowledge of stoichiometry and percentage yield to solve problems related to percentage purity.

Sample Problem

Finding Percentage Purity

Problem

Iron pyrite, FeS_2, is known as "fool's gold" because it looks similar to gold. Suppose that you have a 13.9 g sample of *impure* iron pyrite. (The sample contains a non-reactive impurity.) You heat the sample in air to produce iron(III) oxide, Fe_2O_3, and sulfur dioxide, SO_2.

$$4FeS_{2(s)} + 11O_{2(g)} \rightarrow 2Fe_2O_{3(s)} + 8SO_{2(g)}$$

If you obtain 8.02 g of iron(III) oxide, what was the percentage of iron pyrite in the original sample? Assume that the reaction proceeds to completion. That is, all the available iron pyrite reacts completely.

What Is Required?

You need to determine the percentage purity of the iron pyrite sample.

What Is Given?

The mass of Fe_2O_3 is 8.02 g. The reaction proceeds to completion. You can assume that sufficient oxygen is present.

Plan Your Strategy

Steps 1–4 Use your stoichiometry problem-solving skills to find the mass of Fe_2S expected to have produced 8.02 g Fe_2O_3.

Step 5 Determine percentage purity of the Fe_2S using the following formula:

$$\frac{\text{theoretical mass (g)}}{\text{sample size (g)}} \times 100\%$$

Continued ...

Continued ...
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Act on Your Strategy

1 $$4FeS_{2(s)} + 11O_{2(g)} \longrightarrow 2Fe_2O_{3(s)} + 8SO_{2(g)}$$

0.100 mol ← 0.0502 mol ←

unknown ratio　　known ratio

3 $$\frac{n \text{ mol } FeS_2}{0.0502 \text{ mol } Fe_2O_3} = \frac{4 \text{ mol } FeS_2}{2 \text{ mol } Fe_2O_3}$$

$$(\cancel{0.0502 \text{ mol } Fe_2O_3})\frac{n \text{ mol } FeS_2}{(\cancel{0.0502 \text{ mol } Fe_2O_3})} = \frac{4 \text{ mol } FeS_2}{2 \cancel{\text{ mol } Fe_2O_3}}(0.0502 \cancel{\text{ mol } Fe_2O_3})$$

$$= 0.100 \text{ mol } FeS_2$$

2 $$\frac{8.02 \cancel{\text{ g}} Fe_2O_3}{160 \cancel{\text{ g}}/\text{mol}} = 0.0502 \text{ mol } Fe_2O_3$$

4 $$0.100 \cancel{\text{ mol}} FeS_2 \times 120 \text{ g}/\cancel{\text{mol}} = 12.0 \text{ g } FeS_2$$

12.0 g FeS_2　　8.02 g Fe_2O_3

5 $$\text{Percentage purity} = \frac{\text{Theoretical } m \ FeS_2}{\text{Sample size } FeS_2} \times 100\% = \frac{12.0 \text{ g}}{13.9 \text{ g}} \times 100\% = 86.3\%$$

Therefore, the percentage purity of the iron pyrite is 86.3%.

Check Your Solution

The units are correct. The molar mass of iron pyrite is 3/4 the molar mass of iron(III) oxide. Mutiplying this ratio by the mole ratio of iron pyrite to iron(III) oxide (4/2) and 8 g gives 12 g. The answer is reasonable.

Practice Problems

38. An impure sample of silver nitrate, $AgNO_3$, has a mass 0.340 g. It is dissolved in water and then treated with excess hydrogen chloride, $HCl_{(aq)}$. This results in the formation of a precipitate of silver chloride, AgCl.

$$AgNO_{3(aq)} + HCl_{(aq)} \rightarrow AgCl_{(s)} + HNO_{3(aq)}$$

The silver chloride is filtered, and any remaining hydrogen chloride is washed away. Then the silver chloride is dried. If the mass of the dry silver chloride is measured to be 0.213 g, what mass of silver nitrate was contained in the original (impure) sample?

39. Copper metal is mined as one of several copper-containing ores. One of these ores contains copper in the form of malachite. Malachite exists as a double salt, $Cu(OH)_2 \cdot CuCO_3$. It can be thermally decomposed at 200°C to yield copper(II) oxide, carbon dioxide gas, and water vapour.

$$Cu(OH)_2 \cdot CuCO_{3(s)} \rightarrow 2CuO_{(s)} + CO_{2(g)} + H_2O_{(g)}$$

Continued ...

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(a) 5.000 kg of malachite ore, containing 5.20% malachite, $Cu(OH)_2{\cdot}CuCO_3$, is thermally decomposed. Calculate the mass of copper(II) oxide that is formed. Assume 100% reaction.

(b) Suppose that the reaction had a 78.0% yield, due to incomplete decomposition. How many grams of CuO would be produced?

40. Ethylene oxide, C_2H_4O, is a multi-purpose industrial chemical used, among other things, as a rocket propellant. It can be prepared by reacting ethylene bromohydrin, C_2H_5OBr, with sodium hydroxide.

$$C_2H_5OBr + NaOH \rightarrow C_2H_4O + NaBr + H_2O$$

If this reaction proceeds with an 89% yield, what mass of C_2H_4O can be obtained when 3.61×10^{23} molecules of C_2H_5OBr react with excess sodium hydroxide?

Section Wrap-up

In this section, you have learned how the amount of products formed by experiment relates to the theoretical yield predicted by stoichiometry. You have learned about many factors that affect actual yield, including the nature of the reaction, experimental design and execution, and the purity of the reactants. Usually, when you are performing an experiment in a laboratory, you want to maximize your percentage yield. To do this, you need to be careful not to contaminate your reactants or lose any products. Either might affect your actual yield.

Section Review

1 K/U When calculating the percentage yield of a reaction, what units should you use: grams, moles, or number of particles? Explain.

2 I Methyl salicylate, otherwise known as oil of wintergreen, is produced by the wintergreen plant. It can also be synthesized by heating salicylic acid, $C_7H_6O_3$, with methanol, CH_3OH.

$$C_7H_6O_{3(s)} + CH_3OH_{(\ell)} \rightarrow C_8H_8O_{3(\ell)} + H_2O_{(\ell)}$$

A chemist reacts 3.50 g of salicylic acid with excess methanol. She calculates the theoretical yield of methyl salicylate to be 3.86 g. If 2.84 g of methyl salicylate are recovered, what is the percentage yield of the reaction?

3 C Unbeknownst to a chemist, the limiting reactant in a certain chemical reaction is impure. How will this affect the percentage yield of the reaction? Explain.

4 I You have a sample of copper that is impure, and you wish to determine its purity. You have some silver nitrate, $AgNO_3$, at your disposal. You also have some copper that you know is 100% pure.

(a) Design an experiment to determine the purity of the copper sample.

(b) Even with pure copper, the reaction may not proceed with 100% yield. How will you address this issue?

CHAPTER 7 Review

Reflecting on Chapter 7

Summarize this chapter in the format of your choice. Here are a few ideas to use as guidelines:

- Use the coefficients of a balanced chemical equation to determine the mole ratios between reactants and products.
- Predict quantities required or produced in a chemical reaction.
- Calculate the limiting reactant in cases where the amount of various reactants was given.
- Calculate the percentage yield of a chemical reaction based on the amount of product(s) obtained relative to what was predicted by stoichiometry.
- Use the percentage yield of a reaction to predict the amount of product(s) formed.
- Determine the percentage purity of a reactant based on the actual yield of a reaction.

Reviewing Key Terms

For each of the following terms, write a sentence that shows your understanding of its meaning.

actual yield	competing reaction
excess reactant	limiting reactant
mole ratios	percentage purity
percentage yield	stoichiometric amounts
stoichiometric coefficients	stoichiometry
	theoretical yield

Knowledge/Understanding

1. Explain the different interpretations of the coefficients in a balanced chemical equation.
2. Why is a *balanced* chemical equation needed for stoichiometric calculations?
3. In what cases would it not be necessary to determine the limiting reactant before beginning any stoichiometric calculations?
4. Why was the concept of percentage yield introduced?
5. A student is trying to determine the mass of aluminum oxide that is produced when aluminum reacts with excess oxygen.
$$4Al_{(s)} + 3O_{2(g)} \rightarrow 2Al_2O_{3(s)}$$
The student states that 4 g of aluminum reacts with 3 g of oxygen to produce 2 g of aluminum oxide. Is the student's reasoning correct? Explain your answer.

Inquiry

6. A freshly exposed aluminum surface reacts with oxygen to form a tough coating of aluminum oxide. The aluminum oxide protects the metal from further corrosion.
$$4Al_{(s)} + 3O_{2(g)} \rightarrow 2Al_2O_{3(s)}$$
How many grams of oxygen are needed to react with 0.400 mol of aluminum?
7. Calcium metal reacts with chlorine gas to produce calcium chloride.
$$Ca_{(s)} + Cl_{2(g)} \rightarrow CaCl_{2(s)}$$
How many formula units of $CaCl_2$ are expected from 5.3 g of calcium and excess chlorine?
8. Propane is a gas at room temperature, but it exists as a liquid under pressure in a propane tank. It reacts with oxygen in the air to form carbon dioxide and water vapour.
$$C_3H_{8(\ell)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$$
What mass of carbon dioxide gas is expected when 97.5 g of propane reacts with sufficient oxygen?
9. Powdered zinc and sulfur react in an extremely rapid, exothermic reaction. The zinc sulfide that is formed can be used in the phosphor coating on the inside of a television tube.
$$Zn_{(s)} + S_{(s)} \rightarrow ZnS_{(s)}$$
A 6.00 g sample of Zn is allowed to react with 3.35 g of S.
   (a) Determine the limiting reactant.
   (b) Calculate the mass of ZnS expected.
   (c) How many grams of the excess reactant will remain after the reaction?
10. Titanium(IV) chloride reacts violently with water vapour to produce titanium(IV) oxide and hydrogen chloride gas. Titanium(IV) oxide, when finely powdered, is extensively used in paint as a white pigment.
$$TiCl_{4(s)} + H_2O_{(\ell)} \rightarrow TiO_{2(s)} + 4HCl_{(g)}$$
The reaction has been used to create smoke screens. In moist air, the $TiCl_4$ reacts to produce a thick smoke of suspended TiO_2 particles. What mass of TiO_2 can be expected when 85.6 g of $TiCl_4$ is reacted with excess water vapour?

11. Silver reacts with hydrogen sulfide gas, which is present in the air. (Hydrogen sulfide has the odour of rotten eggs.) The silver sulfide, Ag_2S, that is produced forms a black tarnish on the silver.

$$4Ag_{(s)} + 2H_2S_{(g)} + O_{2(g)} \rightarrow 2Ag_2S_{(s)} + 2H_2O_{(g)}$$

How many grams of silver sulfide are formed when 1.90 g of silver reacts with 0.280 g of hydrogen sulfide and 0.160 g of oxygen?

12. 20.8 g of calcium phosphate, $Ca_3(PO_4)_2$, 13.3 g of silicon dioxide, SiO_2, and 3.90 g of carbon react according to the following equation:

$$2Ca_3(PO_4)_{2(s)} + 6SiO_{2(s)} + 10C_{(s)} \rightarrow P_{4(s)} + 6CaSiO_{3(s)} + 10CO_{(g)}$$

Determine the mass of calcium silicate, $CaSiO_3$, that is produced.

13. 1.56 g of As_2S_3, 0.140 g of H_2O, 1.23 g of HNO_3, and 3.50 g of $NaNO_3$ are reacted according to the equation below:

$$3As_2S_{3(s)} + 4H_2O_{(\ell)} + 10HNO_{3(aq)} + 18NaNO_{3(aq)} \rightarrow 9Na_2SO_{4(aq)} + 6H_3AsO_{4(aq)} + 28NO_{(g)}$$

What mass of H_3AsO_4 is produced?

14. 2.85×10^2 g of pentane, C_5H_{12}, reacts with 3.00 g of oxygen gas, according to the following equation:

$$C_5H_{12(\ell)} + 8O_{2(g)} \rightarrow 5CO_{2(g)} + 6H_2O_{(\ell)}$$

What mass of carbon dioxide gas, is produced?

15. Silica (also called silicon dioxide), along with other silicates, makes up about 95% of Earth's crust—the outermost layer of rocks and soil. Silicon dioxide is also used to manufacture transistors. Silica reacts with hydrofluoric acid to produce silicon tetrafluoride and water vapour.

$$SiO_{2(s)} + 4HF_{(aq)} \rightarrow SiF_{4(g)} + 2H_2O_{(g)}$$

(a) 12.2 g of SiO_2 is reacted with a small excess of HF. What is the theoretical yield, in grams, of H_2O?

(b) If the actual yield of water is 2.50 g, what is the percentage yield of the reaction?

(c) Assuming the yield obtained in part (b), what mass of SiF_4 is formed?

16. An impure sample of barium chloride, $BaCl_2$, with a mass of 4.36 g, is added to an aqueous solution of sodium sulfate, Na_2SO_4.

$$BaCl_{2(s)} + Na_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + 2NaCl_{(aq)}$$

After the reaction is complete, the solid barium sulfate, $BaSO_4$, is filtered and dried. Its mass is found to be 2.62 g. What is the percentage purity of the original barium chloride?

17. Benzene reacts with bromine to form bromobenzene, C_6H_5Br.

$$C_6H_{6(\ell)} + Br_{2(\ell)} \rightarrow C_6H_5Br_{(\ell)} + HBr_{(g)}$$

(a) What is the maximum amount of C_6H_5Br that can be formed from the reaction of 7.50 g of C_6H_6 with excess Br_2?

(b) A competing reaction is the formation of dibromobenzene, $C_6H_4Br_2$.

$$C_6H_{6(\ell)} + 2Br_{2(\ell)} \rightarrow C_6H_4Br_{2(\ell)} + 2HBr_{(g)}$$

If 1.25 g of $C_6H_4Br_2$ was formed by the competing reaction, how much C_6H_6 was *not* converted to C_6H_5Br?

(c) Based on your answer to part (b), what was the actual yield of C_6H_5Br? Assume that all the C_6H_5Br that formed was collected.

(d) Calculate the percentage yield of C_6H_5Br.

18. Refer to Practice Problem 39. Design an experiment to determine the mole to mole ratio of pure malachite to copper(II) oxide. Include an outline of the procedure and any safety precautions. Clearly indicate which data need to be recorded.

19. A chemist wishes to prepare a compound called compound E. The molar mass of compound E is 100 g/mol. The synthesis requires four consecutive reactions, each with a yield of 60%.

$A \rightarrow B$ $\quad$ $C \rightarrow D$

$B \rightarrow C$ $\quad$ $D \rightarrow E$

(a) The chemist begins the synthesis with 50 g of starting material, called compound A. If the molar mass of compound A is 200 g/mol, how many grams of compound E will be produced?

(b) How many grams of compound A are needed to produce 70 g of compound E?

Communication

20. Develop a new analogy for the concept of limiting and excess reactant.

21. Examine the balanced chemical "equation"

$$2A + B \rightarrow 3C + D$$

Using a concept map, explain how to calculate the number of grams of C that can be obtained when a given mass of A reacts with a certain number of molecules of B. Assume that you know the molar mass of A and C. Include proper units. For simplicity, assume that A is limiting, but don't forget to show how to determine the limiting reactant.

22. Assume that your friend has missed several chemistry classes and that she has asked you to help her prepare for a stoichiometry test. Unfortunately, because of other commitments, you do not have time to meet face to face. You agree to email your friend a set of point-form instructions on how to solve stoichiometry problems, including those that involve a limiting reactant. She also needs to understand the concept of percentage yield. Write the text of this email. Assume that your friend has a good understanding of the mole concept.

Making Connections

23. How many grams of air are required for an automobile to travel from Thunder Bay, Ontario, to Smooth Rock Falls, Ontario? This is a distance of 670 km. Assume the following:

- Gasoline is pure octane, C_8H_{18}. (Gasoline is actually a mixture of hydrocarbons.)
- The average fuel consumption is 10 L per 100 km.
- Air has a density of 1.21 g/L.
- Air is 21% O_2 (v/v).
- 1.00 mol of any gas occupies 24 L at 20°C and 100 kPa.
- The density of the gasoline is 0.703 g/mL.

The balanced chemical equation for the complete combustion of octane is

$$2C_8H_{18(\ell)} + 25O_{2(g)} \rightarrow 16CO_{2(g)} + 18H_2O_{(g)}$$

24. You must remove mercury ions present as mercury(II) nitrate in the waste water of an industrial facility. You have decided to use sodium sulfide in the reaction below. Write a short essay that addresses the following points. Include a well-organized set of calculations where appropriate.

$$Hg(NO_3)_{2(aq)} + Na_2S_{(aq)} \rightarrow HgS_{(s)} + 2NaNO_{3(aq)}$$

(a) Explain why the chemical reaction above can be used to remove mercury ions from the waste water. What laboratory technique must be used in order that this reaction is as effective as possible for removing mercury from the waste stream?

(b) Why is mercury(II) sulfide less of an environmental concern than mercury(II) nitrate?

(c) What assumptions are being made regarding the toxicity of sodium sulfide and sodium nitrate relative to either mercury nitrate or mercury sulfide?

(d) Every litre of waste water contains approximately 0.03 g of $Hg(NO_3)_2$. How many kg of Na_2S will be required to remove the soluble mercury ions from 10 000 L of waste water?

(e) What factors would a company need to consider in adopting any method of cleaning its wastewater?

Answers to Practice Problems and Short Answers to Section Review Questions

Practice Problems: **1.(a)** 2:1:2 **(b)** 50 **(c)** 4956 **(d)** 1.20×10^{24} **2.(a)** 2 **(b)** 150 **(c)** 1.806×10^{24} **(d)** 1.204×10^{24} **3.(a)** 3.4×10^{25} **(b)** 6.7×10^{25} **4.** 7.5 mol **5.(a)** 1.8 mol **(b)** 37.5 mol **6.(a)** 48.7 mol **(b)** 1.20 mol **7.(a)** 8.3×10^{24} **(b)** 4.2×10^{24} **8.(a)** 7.47 mol **(b)** 7.19 mol **9.(a)** 4.68×10^{-2} mol **(b)** 0.187 mol **10.(a)** 0.708 mol **(b)** 1.06 mol **11.** 9.32 g **12.** 137 g **13.** 4.65 g **14.** 0.814 g **15.** 97.2 g **16.** 2.31×10^{-2} g **17.** 37.6 g **18.** 20.7 g **19.(a)** 124 g **(b)** 1.14×10^{24} **20.(a)** 120 g **(b)** 1.49×10^{21} **21.(a)** 2.39×10^{22} **(b)** 1.45×10^{24} **(c)** 1.21×10^{22} **22.(a)** 1.50×10^{24} **(b)** 357 g **23.** $CuCl_2$ **24.** CaF_2 **25.** C_3H_6 **26.** HCl **27.(a)** ClO_2 **(b)** 74.1 g **(c)** 1.06×10^{23} **28.(a)** H_2O_2 **(b)** 63.5 g **(c)** 104 g **29.** 4.23×10^4 g **30.(a)** 0.446 g **(b)** 4.80×10^{21} **(c)** F_2, 24.0 g **31.(a)** 74.4 g **(b)** 63.6% **32.(a)** 31.3 g **(b)** 95.2% **33.** 26.7% **34.** 14.1 g **35.(a)** 15.0 g **(b)** 10.5 g **(c)** 12.8 g **36.** 5.63 g **37.** 129 g **38.** 0.252 g **39.(a)** 187 g **(b)** 146 g **40.** 23.5 g

Section Review: **7.1:** **4.(a)** $S + O_2 \rightarrow SO_2$ **(b)** $2S + 3O_2 \rightarrow 2SO_3$ **(c)** 1.5 mol **(d)** 48.0 g **5.(a)** 1:5:3:4 **(b)** 2.50 mol **(c)** 6.02×10^{24} **(d)** 9.00 mol **6.(a)** 49.9 g **(b)** 4.48×10^{22} **7.(a)** $Pb(NO_3)_{2(aq)} + 2NaI_{(aq)} \rightarrow PbI_{2(s)} + 2NaNO_{3(aq)}$ **(b)** 1.03 g **7.2:** **2.(a)** oxygen **3.** 18.1 g **4.** 8.04 g **5.** 5.34 g **6.** 22.2 g **7.(a)** $Zn_{(s)} + CuCl_{2(aq)} \rightarrow ZnCl_{2(aq)} + Cu_{(s)}$ **(b)** zinc gone **(c)** zinc **(d)** less than 1.52 g Zn **7.3:** **2.** 73.6%