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.

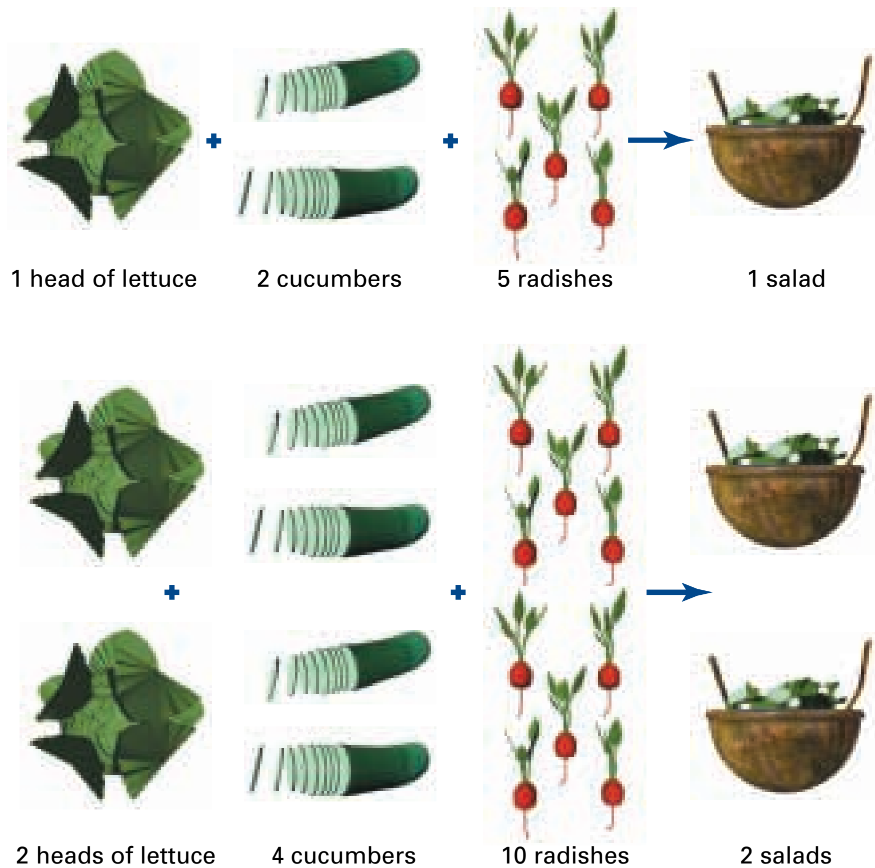

Figure 7.1 **A salad analogy showing how equations can be multiplied**

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.

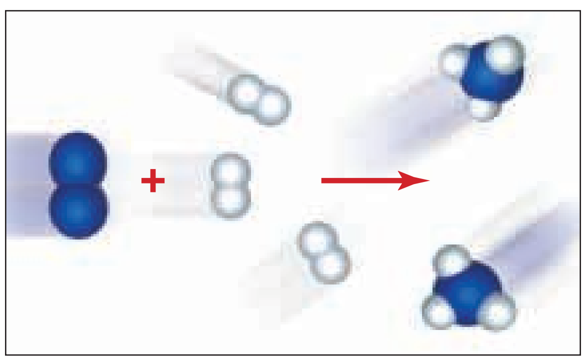

Figure 7.3 The reaction of nitrogen gas with hydrogen gas.

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.

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

(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)?

(c) How many molecules of water are formed when 2478 molecules of O_2 react with H_2?

(d) How many molecules of H_2 are required to react completely with 6.02×10^{23} molecules of O_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.

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

(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)?

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

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

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.

unknown ratio **known ratio**

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

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

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

Try the following Practice Problems to work with mole ratios.

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)?

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.

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

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.

(b) How many moles of V are needed to react with 5.39 mol of O_2 to produce V_2O_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:

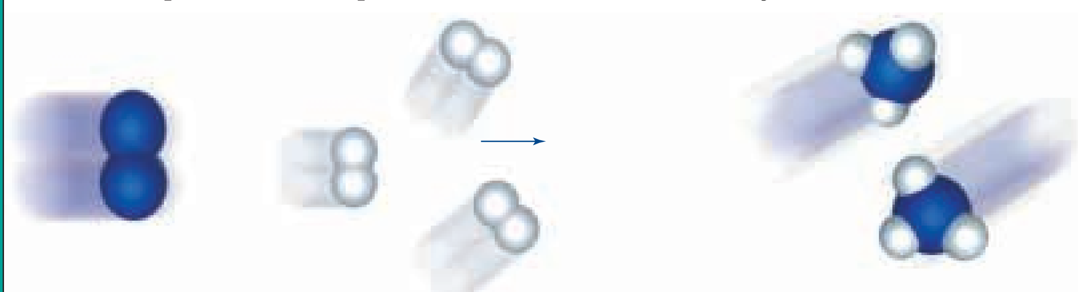
$$1 \text{ mol } N_2 \times 28.0 \text{ g/mol } N_2 = 28.0 \text{ g } N_2$$
$$3 \text{ mol } H_2 \times 2.02 \text{ g/mol } H_2 = 6.1 \text{ g } H_2$$
$$2 \text{ mol } NH_3 \times 17.0 \text{ g/mol } NH_3 = 34.1 \text{ g } NH_3$$

CHECKPOINT

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

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.

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?

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.**

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

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

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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.

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

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

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?

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.

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?

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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.

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

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

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.

Electronic Learning Partner

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

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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.

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

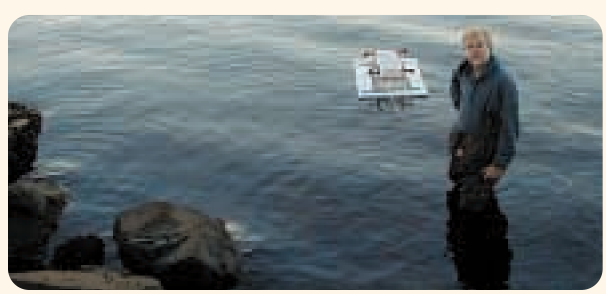
How many grams of magnesium oxide are required to react completely with 33.5 g of phosphoric acid?

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.

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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.

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

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.

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?

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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?