Stoichiometry

Wade Baxter, Ph.D.

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AUTHOR
Wade Baxter, Ph.D.

EDITORS
Donald Calbreath, Ph.D.
Max Helix

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CHAPTER 1 Stoichiometry

CHAPTER OUTLINE

Artists use paints to create beautiful images and those paints come in a wide variety of distinctive colors. However, an artist will frequently mix paints to get just the right hue for a certain subject that he or she is painting. Mixing blue and green in different ratios will give many different shades of turquoise, for example. Chemists also have to be concerned with the ratios of reactants that are required for a particular chemical reaction. In this chapter, you will learn how it is the specific ratio of moles of reactants that governs reactions and determines the amount of products that can be formed in a specific reaction. You will combine your knowledge of mole relationships with your ability to balance chemical equations in order to gain a complete understanding of the quantitative nature of chemical reactions.

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1.1 Mole Ratios

Lesson Objectives

- Relate balanced chemical equations to everyday analogies, such as a recipe.
- Define stoichiometry.
- Use mole ratios to convert between amounts of substances in a chemical reaction.

Lesson Vocabulary

- mole ratio
- stoichiometry

Check Your Understanding

Recalling Prior Knowledge

- Why is it necessary to balance chemical equations?
- What do the coefficients of a balanced chemical equation represent?

Chemical equations are balanced in order to satisfy the law of conservation of mass. The mass of each element in a chemical reaction must be conserved. A balanced equation is required for calculations that involve the quantities of reactants used and products generated in a chemical reaction.

Everyday Stoichiometry

In the last chapter, *Chemical Reactions*, you learned about chemical equations and the techniques used to balance them. Chemists use balanced equations to obtain quantitative information about chemical reactions. Before we look at a chemical reaction, let's reconsider the equation for the ideal ham sandwich.

As a reminder, our ham sandwich is composed of 2 slices of ham (H), 1 slice of cheese (C), 1 slice of tomato (T), 5 pickles (P), and 2 slices of bread (B). The equation for our sandwich is shown below:

$$2H + C + T + 5P + 2B \rightarrow H_2CTP_5B_2$$

Now let us suppose that you are having some friends over and need to make five ham sandwiches. How much of each sandwich ingredient do you need? You would take the number of each ingredient required for one sandwich (its coefficient in the above equation) and multiply by five. Using ham and cheese as examples and using a conversion factor, we can write:

$$5\ H_2CTP_5B_2 \times \frac{2\ H}{1\ H_2CTP_5B_2} = 10\ H$$
$$5\ H_2CTP_5B_2 \times \frac{1\ C}{1\ H_2CTP_5B_2} = 5\ C$$

The conversion factors contain the coefficient of each specific ingredient as the numerator and the formula of one sandwich as the denominator. The result is what you would expect. In order to make five ham sandwiches, you would need 10 slices of ham and 5 slices of cheese.

This type of calculation demonstrates the use of stoichiometry. **Stoichiometry** *is the calculation of amounts of substances in a chemical reaction from the balanced equation*. The sample problem below is another stoichiometry problem involving ingredients of the ideal ham sandwich.

Sample Problem 12.1: Ham Sandwich Stoichiometry

Kim looks in the refrigerator and finds that she has 8 slices of ham. In order to make as many sandwiches as possible, how many pickles does she need? Use the equation from the text.

Step 1: List the known quantities and plan the problem.

Known

- have 8 ham slices (H)
- 2 H = 5 P (conversion factor)

Unknown

- How many pickles (P) needed?

The coefficients for the two reactants (ingredients) are used to make a conversion factor between ham slices and pickles.

Step 2: Solve.

$$8\ H \times \frac{5\ P}{2\ H} = 20\ P$$

Since 5 pickles combine with 2 ham slices in each sandwich, 20 pickles are needed to fully combine with 8 ham slices.

Step 3: Think about your result.

The 8 ham slices will make 4 ham sandwiches. Since there are 5 pickles per sandwich, 20 pickles would be used in these 4 sandwiches.

Practice Problems

1. How many sandwiches can be made from 30 pickles, assuming that you have enough of the rest of the ingredients?
2. You have 7 slices of cheese. How much of each of the other ingredients is required to make the maximum number of sandwiches possible? How many sandwiches could be made?

Watch this short video for an introduction to the concept of stoichiometry: http://www.youtube.com/watch?v=xGBtz-ihRbA.

MEDIA

Click image to the left for more content.

Mole Ratios

Stoichiometry problems can be characterized by two things: (1) the information *given* in the problem, and (2) the information that is to be solved for, referred to as the *unknown*. The given and the unknown may both be reactants, both be products, or one may be a reactant while the other is a product. The amounts of the substances can be expressed in moles. However, in a laboratory situation, it is common to determine the amount of a substance by finding its mass in grams. The amount of a liquid or gaseous substance may also be expressed by its volume. In this lesson, we will focus on the type of problem where both the given and the unknown quantities are expressed in moles.

Given substance (mol) $\rightarrow$ Unknown substance (mol)

In the following lesson, "Stoichiometric Calculations," we will expand our analysis of stoichiometry to include mass-based and volume-based problems.

Chemical equations express the relative amounts of reactants and products in a reaction. The coefficients of a balanced equation can represent either the number of molecules or the number of moles of each substance. The production of ammonia (NH_3) from nitrogen and hydrogen gases is an important industrial reaction called the Haber process (**Figure** 1.1), after German chemist Fritz Haber.

FIGURE 1.1

Nitrogen and hydrogen gases are combined under careful conditions to form ammonia in a reaction called the Haber process. Ammonia is a very important industrial chemical and is produced in very large quantities, mostly for fertilizers.

The balanced equation for the Haber process is:

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

The reaction can be analyzed in several ways, as shown below (**Figure** 1.2).

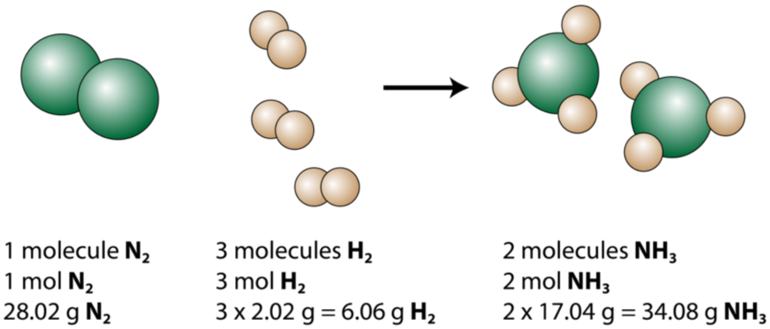


FIGURE 1.2
This representation of the production of ammonia from nitrogen and hydrogen shows several ways to interpret the quantitative information given by a balanced chemical equation.

1 molecule of nitrogen reacts with 3 molecules of nitrogen to form 2 molecules of ammonia. These are the smallest possible relative amounts of the reactants and products. To consider larger relative amounts, each coefficient can be multiplied by the same number. For example, 10 molecules of nitrogen would react with 30 molecules of hydrogen to produce 20 molecules of ammonia.

As you have learned, the most useful quantity for counting very tiny particles is the mole. If each coefficient is multiplied by a mole, the balanced chemical equation tells us that 1 mole of nitrogen reacts with 3 moles of hydrogen to produce 2 moles of ammonia. This is the conventional way to interpret any balanced chemical equation.

Notice that the number of moles of reactants is not conserved when they are converted to products. In the production of ammonia, 4 moles of reactant molecules are converted into 2 moles of product molecules. However, if each mole quantity is converted to grams by using the molar mass, we can see that the law of conservation of mass is followed. 1 mol of nitrogen has a mass of 28.02 g, 3 mol of hydrogen has a mass of 6.06 g, and 2 mol of ammonia has a mass of 34.08 g.

$$28.02 \text{ g } N_2 + 6.06 \text{ g } H_2 \rightarrow 34.08 \text{ g } NH_3$$

Mass and the total number of each atom must be conserved in any chemical reaction. The number of molecules is not necessarily conserved.

A **mole ratio** *is a conversion factor that relates the amounts in moles of any two substances in a chemical reaction*. The numbers in a conversion factor come from the coefficients of the balanced chemical equation. The following six mole ratios can be written for the reaction above.

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

In a mole ratio problem, the given amount, expressed in moles, is written first. The appropriate conversion factor is chosen in order to convert from moles of the given substance to moles of the unknown substance.

Sample Problem 12.2: Mole Ratio

How many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen?

Step 1: List the known quantities and plan the problem.

Known

- given: H_2 = 4.20 mol

Unknown

- mol of NH_3

The conversion is from mol $H_2 \rightarrow$ mol NH_3. The problem states that there is an excess of nitrogen, so we do not need to be concerned with any mole ratio involving N_2. Choose the conversion factor that has moles of NH_3 in the numerator and moles of H_2 in the denominator so that we are left with the desired quantity after canceling units.

Step 2: Solve.

$$4.20 \text{ mol } H_2 \times \frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2} = 2.80 \text{ mol } NH_3$$

The reaction of 4.20 mol of hydrogen with excess nitrogen produces 2.80 mol of ammonia.

Step 3: Think about your result.

The result corresponds to the 3:2 ratio of hydrogen to ammonia from the balanced equation. Note that a mole ratio represents exact quantities, since the coefficients in a balanced equation are considered to have an infinite number of significant figures. The number of significant figures in the result is determined directly from the number of significant figures in the given quantity. In this case, both have three significant figures.

Practice Problems

3. How many moles of hydrogen are required to fully react with 0.26 moles of nitrogen?
4. A reaction produces 3.12 moles of ammonia. How many moles of nitrogen and hydrogen were reacted?

Watch an animation showing the mole ratio of hydrogen to oxygen when they react to form water at http://www.dlt.ncssm.edu/core/Chapter6-Stoichiometry/Chapter6-Animations/OneLiterH2O.html.

Practice mole ratios using both sandwiches and real chemical reactions using the simulation at http://phet.colorado.edu/en/simulation/reactants-products-and-leftovers. Several activities associated with this lesson are available by scrolling down.

Lesson Summary

- A balanced chemical equation provides the same information as a recipe. Recipes can be manipulated to account for different amounts of various ingredients.
- Stoichiometry is the branch of chemistry that involves relationships between the amounts of substances in a balanced equation. The coefficients represent the moles of each reactant and product in the reaction.
- Mole ratios can be constructed between any two of the substances in a balanced equation. Moles of a given substance can be converted to moles of an unknown substance with the appropriate mole ratio.

Lesson Review Questions

Reviewing Concepts

1. The balanced equation for the combustion of ethene (C_2H_6) is given below:
   a. Interpret the equation in terms of the number of molecules of each substance that are involved in the reaction.
   b. Describe the reaction again in terms of moles of each substance.
   c. Determine the mass of each reactant and product in the equation, and demonstrate that the law of conservation of mass is obeyed.
2. For the reaction given in question 1, write the following mole ratios:
   a. ethene to oxygen
   b. carbon dioxide to water
   c. ethene to carbon dioxide
3. Write all of the possible mole ratios for the equation below, in which calcium reacts with nitrogen to form calcium nitride:

Problems

4. A certain automobile contains 4 tires, 2 headlights, 1 steering wheel, and 6 spark plugs. How many of each part will be required to build 17 of these cars?
5. The reaction of sodium with water produces sodium hydroxide and hydrogen gas: $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$
   a. How many moles of water are required to react with 0.89 moles of sodium?
   b. The complete reaction of 1.92 moles of sodium will produce how many moles of hydrogen?
   c. If 0.00482 moles of H_2 were produced by this reaction, how many moles of NaOH were also produced?
6. Ammonia will react with nitrogen monoxide to produce nitrogen gas and water vapor. $4NH_3(g) + 6NO(g) \rightarrow 5N_2(g) + 6H_2O(g)$
   a. How many moles of N_2 will be produced by the complete reaction of 3.20 moles of NH_3?
   b. In order to completely react with 0.810 moles of NO, how many moles of NH_3 are needed?
   c. How many moles of NH_3 were consumed in a reaction that produced 9.04 moles of H_2O?
   d. You have 2.28 moles of NH_3. How many moles of NO are needed to react with this amount of ammonia? How many moles of each of the products will be formed?
7. The reaction of aluminum metal with an aqueous solution of silver nitrate produces aqueous aluminum nitrate and solid silver metal.
   a. Write the balanced chemical equation for this reaction.
   b. If 0.612 moles of aluminum are reacted, how many moles of silver will be produced?

Further Reading / Supplemental Links

- Mole Ratios in Chemical Equations, (http://www.wisc-online.com/objects/ViewObject.aspx?ID=GCH7304)
- Chemistry: Mole-Mole Ratios, (http://www.algebralab.org/practice/practice.aspx?file=Algebra_StoichiometryMoletoMoleConversions.xml)

Points to Consider

In the laboratory, there is no way for a chemist to directly measure out moles of a substance. He or she, instead, must determine amounts either by mass or by volume. Therefore, it is important to be able to perform stoichiometric calculations with mass and volume.

- How is mass converted to moles?
- How are moles converted to mass?
- How many conversion factors are required to convert from the mass of a given substance to the mass of an unknown substance?

1.2 Stoichiometric Calculations

Lesson Objectives

- Calculate the amount in moles of a reactant or product from the mass of another reactant or product. Calculate the mass of a reactant or product from the moles of another reactant or product.
- Calculate the mass of a reactant or product from the mass of another reactant or product.
- Create volume ratios from a balanced chemical equation.
- Use volume ratios and other stoichiometric principles to solve problems involving mass, molar amounts, or volumes of gases.

Check Your Understanding

Recalling Prior Knowledge

- How can a balanced chemical equation be used to construct mole ratios between substances?
- How is a mole ratio used to convert from moles of one reactant or product to moles of another?

The mole ratio is the essence of ideal stoichiometry. The mole ratio tells us the quantitative relationship between reactants and products under ideal conditions, in which all reactants are completely converted into products. In the laboratory, most reactions are not completely ideal. Reactions may not proceed 100% to completion, or a given set of reactants may also undergo side reactions that lead to different products. However, theoretical stoichiometric calculations are important because they allow chemists to know the maximum possible amount of product that can be generated by a reaction from a given amount of each reactant.

Ideal Stoichiometry

Solving any stoichiometric calculation starts with a balanced chemical equation. As we saw in the last lesson, "Mole Ratios," the coefficients of the balanced equation are the basis for the mole ratio between any pair of reactants and/or products. The following flowchart (**Figure** 1.3) shows that the conversion from a given substance in moles to moles of an unknown substance involves multiplying by the relevant mole ratio.

In this lesson, you will expand your understanding of stoichiometry to include the amounts of substances that are measured either by mass or by volume.

Mass-Based Stoichiometry

While the mole ratio is ever-present in all stoichiometry calculations, amounts of substances in the laboratory are most often measured by mass. Therefore, we need to use mole-mass calculations in combination with mole ratios to

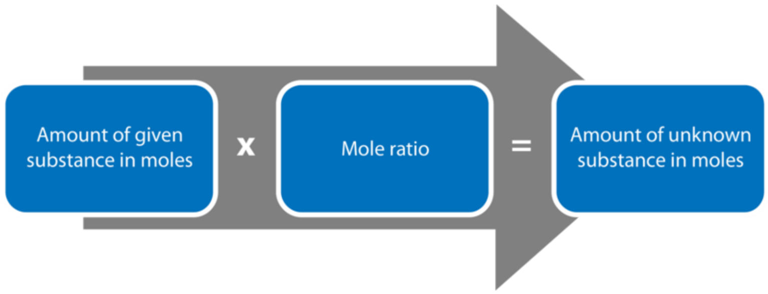


FIGURE 1.3
This flowchart shows how a mole ratio is used in a stoichiometric conversion problem.

solve several different types of mass-based stoichiometry problems.

Mass to Moles Problems

In this type of problem, the mass of one substance is given, usually in grams. This value is then used to determine the amount in moles of another substance that will either react with or be produced from the given substance.

mass of *given* $\rightarrow$ moles of *given* $\rightarrow$ moles of *unknown*

The mass of the given substance is converted into moles by using the molar mass of that substance, which can be calculated from the atomic masses found on a periodic table. Then, the moles of the given substance are converted into moles of the unknown by using the mole ratio from the balanced chemical equation.

Sample Problem 12.3: Mass-Mole Stoichiometry

Tin metal reacts with hydrogen fluoride to produce tin(II) fluoride and hydrogen gas according to the following balanced equation:

$Sn(s) + 2HF(g) \rightarrow SnF_2(s) + H_2(g)$

How many moles of hydrogen fluoride are required to react completely with 75.0 g of tin?

Step 1: List the known quantities and plan the problem.

Known

- given: 75.0 g Sn
- molar mass of Sn = 118.69 g/mol
- 1 mol Sn = 2 mol HF (mole ratio)

Unknown

- mol HF

Use the molar mass of Sn to convert the given mass of Sn to moles. Then use the mole ratio to convert from mol Sn to mol HF. This will be done in a single two-step calculation.

g Sn $\rightarrow$ mol Sn $\rightarrow$ mol HF

Step 2: Solve.

$$75.0 \text{ g Sn} \times \frac{1 \text{ mol Sn}}{118.69 \text{ g Sn}} \times \frac{2 \text{ mol HF}}{1 \text{ mol Sn}} = 1.26 \text{ mol HF}$$

Step 3: Think about your result.

The mass of tin is less than one mole, but the 1:2 ratio means that more than one mole of HF is required for the reaction. The answer has three significant figures because the given mass has three significant figures.

Practice Problem

1. Silver oxide is used in small batteries called button batteries (**Figure** 1.4). It decomposes upon heating to form silver metal and oxygen gas.

$2Ag_2O(s) \rightarrow 4Ag(s) + O_2(g)$

How many moles of silver are produced by the decomposition of 4.85 g of Ag_2O?

Moles to Mass Problems

In this type of problem, the amount of one substance is given in moles. From this, you can determine the mass of another substance that will either react with or be produced from the given substance.

moles of *given* → moles of *unknown* → mass of *unknown*

The moles of the given substance are first converted into moles of the unknown by using the mole ratio from the balanced chemical equation. Then, the moles of the unknown are converted into a mass (in grams) by using its molar mass, which again can be calculated from the atomic masses given on the periodic table.

Sample Problem 12.4: Mole-Mass Stoichiometry

Hydrogen sulfide gas burns in oxygen to produce sulfur dioxide and water vapor.

$2H_2S(g) + 3O_2(g) \rightarrow 2SO_2(g) + 2H_2O(g)$

What mass of oxygen gas is consumed in a reaction that produces 4.60 moles of SO_2?

Step 1: List the known quantities and plan the problem.

Known

- given: 4.60 mol SO_2
- 2 mol SO_2 = 3 mol O_2 (mole ratio)
- molar mass of O_2 = 32.00 g/mol

Unknown

- mass O_2 = ? g

Use the mole ratio to convert from mol SO_2 to mol O_2. Then convert mol O_2 to grams. This will be done in a single two-step calculation.

mol SO_2 → mol O_2 → g O_2

Step 2: Solve.

$$4.60 \text{ mol } SO_2 \times \frac{3 \text{ mol } O_2}{2 \text{ mol } SO_2} \times \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} = 221 \text{ g } O_2$$

Step 3: Think about your result.

According to the mole ratio, 6.90 mol O_2 is produced, which has a mass of 221 g. The answer has three significant figures because the given value has three significant figures.

Practice Problem

2. Copper metal reacts with sulfur to form copper(I) sulfide. What mass of copper(I) sulfide is produced by the reaction of 0.528 mol Cu with excess sulfur?

Mass to Mass Problems

Mass-mass calculations are the most practical of all mass-based stoichiometry problems. Moles cannot be measured directly, but masses can be easily measured in the lab for most substances. This type of problem is three steps and is a combination of the two previous types.

mass of *given* → moles of *given* → moles of *unknown* → mass of *unknown*

The mass of the given substance is converted into moles by using its molar mass. Then, the moles of the given substance are converted into moles of the unknown by using the mole ratio from the balanced chemical equation. Finally, the moles of the unknown are converted back to a mass by using its molar mass.

Sample Problem 12.5: Mass-Mass Stoichiometry

Ammonium nitrate decomposes to dinitrogen monoxide and water according to the following equation.

$$NH_4NO_3(s) \rightarrow N_2O(g) + 2H_2O(l)$$

In a certain experiment, 45.7 g of ammonium nitrate is decomposed. Find the mass of each of the products formed.

Step 1: List the known quantities and plan the problem.

Known

- given: 45.7 g NH_4NO_3
- 1 mol NH_4NO_3 = 1 mol N_2O = 2 mol H_2O (mole ratios)
- molar mass of NH_4NO_3 = 80.06 g/mol
- molar mass of N_2O = 44.02 g/mol
- molar mass of H_2O = 18.02 g/mol

Unknown

- mass N_2O = ? g
- mass H_2O = ? g

Perform two separate three-step mass-mass calculations, as shown below.

$g\ NH_4NO_3 \rightarrow mol\ NH_4NO_3 \rightarrow mol\ N_2O \rightarrow g\ N_2O$
$g\ NH_4NO_3 \rightarrow mol\ NH_4NO_3 \rightarrow mol\ H_2O \rightarrow g\ H_2O$

Step 2: Solve.

$$45.7\text{ g }NH_4NO_3 \times \frac{1\text{ mol }NH_4NO_3}{80.06\text{ g }NH_4NO_3} \times \frac{1\text{ mol }N_2O}{1\text{ mol }NH_4NO_3} \times \frac{44.02\text{ g }N_2O}{1\text{ mol }N_2O} = 25.1\text{ g }N_2O$$
$$45.7\text{ g }NH_4NO_3 \times \frac{1\text{ mol }NH_4NO_3}{80.06\text{ g }NH_4NO_3} \times \frac{2\text{ mol }H_2O}{1\text{ mol }NH_4NO_3} \times \frac{18.02\text{ g }H_2O}{1\text{ mol }H_2O} = 20.6\text{ g }H_2O$$

Step 3: Think about your result.

The total mass of the two products is equal to the mass of ammonium nitrate that decomposed, demonstrating the law of conservation of mass. Each answer has three significant figures.

Practice Problem

3. Solid iron(III) hydroxide reacts with sulfuric acid to produce aqueous iron(III) sulfate and water. What mass of sulfuric acid is needed to completely react with 12.72 g of iron(III) hydroxide? What mass of iron(III) sulfate is produced?

Volume-Based Stoichiometry

As you learned in the chapter on *The Mole*, Avogadro's hypothesis states that equal volumes of all gases at the same temperature and pressure contain the same number of gas particles. We also saw that one mole of any gas at standard temperature and pressure (0°C and 1 atm) occupies a volume of 22.4 L. These characteristics make stoichiometry calculations involving gases at STP very straightforward. Consider the following reaction, in which nitrogen and oxygen combine to form nitrogen dioxide.

$N_2(g)$	+	$2O_2(g)$	$\rightarrow$	$2NO_2(g)$
1 molecule		2 molecules		2 molecules
1 mol		2 mol		2 mol
1 volume		2 volumes		2 volumes

Because of Avogadro's hypothesis, we know that mole ratios between substances in a gas-phase reaction are also volume ratios. The six possible volume ratios for the above equation are:

- $\frac{1\text{ volume }N_2}{2\text{ volumes }O_2}$ or $\frac{2\text{ volumes }O_2}{1\text{ volume }N_2}$
- $\frac{1\text{ volume }N_2}{2\text{ volumes }NO_2}$ or $\frac{2\text{ volumes }NO_2}{1\text{ volume }N_2}$
- $\frac{2\text{ volumes }O_2}{2\text{ volumes }NO_2}$ or $\frac{2\text{ volumes }NO_2}{2\text{ volumes }O_2}$

Volume to Volume Problems

The volume ratios above can easily be used when the volume of one gas in a reaction is known and you need to determine the volume of another gas that will either react with or be produced from the first gas. Although pressure and temperature need to be held constant over the course of the reaction for these conversion factors to remain true, the reaction does not need to be run at STP; Avogadro's hypothesis is true regardless of the pressure and temperature being used.

Sample Problem 12.6: Volume-Volume Stoichiometry

The combustion of propane gas produces carbon dioxide and water vapor (**Figure** 1.5).

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

FIGURE 1.4
Silver oxide is used in button batteries such as these watch batteries.

What volume of oxygen is required to completely combust 0.650 L of propane? What volume of carbon dioxide is produced in the reaction?

Step 1: List the known quantities and plan the problem.

Known

- given: 0.650 L C_3H_8
- 1 volume C_3H_8 = 5 volumes O_2
- 1 volume C_3H_8 = 3 volumes CO_2

Unknown

- volume O_2 = ? L
- volume CO_2 = ? L

Step 2: Solve.

$$0.650 \text{ L } C_3H_8 \times \frac{5 \text{ L } O_2}{1 \text{ L } C_3H_8} = 3.25 \text{ L } O_2$$
$$0.650 \text{ L } C_3H_8 \times \frac{3 \text{ L } CO_2}{1 \text{ L } C_3H_8} = 1.95 \text{ L } CO_2$$

Step 3: Think about your result.

Because the coefficients on O_2 and CO_2 are larger than the one in front of C_3H_8, the volumes for those two gases are greater. Note that total volume is not necessarily conserved in a reaction because moles are not necessarily conserved. In this reaction, 6 total volumes of reactants become 7 total volumes of products.

Practice Problem

4. Using the equation for the combustion of propane from Sample Problem 12.6, what volume of water vapor is produced in a reaction that generates 320. mL of carbon dioxide?

Mass to Volume and Volume to Mass Problems

Chemical reactions frequently involve both solid substances whose mass can be measured as well as gases for which measuring the volume is more appropriate. Stoichiometry problems of this type are called either mass-volume or volume-mass problems.

mass of *given* → moles of *given* → moles of *unknown* → volume of *unknown*
volume of *given* → moles of *given* → moles of *unknown* → mass of *unknown*

Because both types of problems involve a conversion from either moles of gas to volume or vice-versa, we can use the molar volume of 22.4 L/mol as a conversion factor if the reaction is run at STP. In a later chapter, *The Behavior of Gases*, we will see how to solve this type of problem when other sets of reaction conditions are used.

Sample Problem 12.7: Mass-Volume Stoichiometry

Aluminum metal reacts rapidly with aqueous sulfuric acid to produce aqueous aluminum sulfate and hydrogen gas.

$$2\text{Al}(s) + 3\text{H}_2\text{SO}_4(aq) \rightarrow \text{Al}_2(\text{SO}_4)_3(aq) + 3\text{H}_2(g)$$

Determine the volume of hydrogen gas produced at STP when a 2.00 g piece of aluminum completely reacts with sulfuric acid.

Step 1: List the known quantities and plan the problem.

Known

- given: 2.00 g Al
- molar mass of Al = 26.98 g/mol
- 2 mol Al = 3 mol H_2

Unknown

- volume H_2 = ?

The grams of aluminum will first be converted to moles. Then the mole ratio will be applied to convert to moles of hydrogen gas. Finally, the molar volume of a gas will be used to convert to liters of hydrogen.

$$\text{g Al} \rightarrow \text{mol Al} \rightarrow \text{mol H}_2 \rightarrow \text{L H}_2$$

Step 2: Solve.

$$2.00 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{3 \text{ mol H}_2}{2 \text{ mol Al}} \times \frac{22.4 \text{ L H}_2}{1 \text{ mol H}_2} = 2.49 \text{ L H}_2$$

Step 3: Think about your result.

The volume result is in liters. For smaller amounts, it may be convenient to convert to milliliters. The answer here has three significant figures. Because the molar volume is a measured quantity of 22.4 L/mol, three is the maximum number of significant figures for this type of problem.

Sample Problem 12.8: Volume-Mass Stoichiometry

Calcium oxide is used to trap the sulfur dioxide that is generated in coal-burning power plants according to the following reaction:

$$2CaO(s) + 2SO_2(g) + O_2(g) \rightarrow 2CaSO_4(s)$$

What mass of calcium oxide is required to react completely with 1.2×10^3 L of sulfur dioxide at STP?

Step 1: List the known quantities and plan the problem.

Known

- given: 1.2×10^3 L SO_2
- 2 mol SO_2 = 2 mol CaO
- molar mass of CaO = 56.08 g/mol

Unknown

- mass CaO = ? g

The volume of SO_2 will first be converted to moles. The mole ratio can then be used, and finally, moles of CaO will be converted to grams.

$$\text{L } SO_2 \rightarrow \text{mol } SO_2 \rightarrow \text{mol CaO} \rightarrow \text{g CaO}$$

Step 2: Solve.

$$1.2 \times 10^3 \text{ L } SO_2 \times \frac{1 \text{ mol } SO_2}{22.4 \text{ L } SO_2} \times \frac{2 \text{ mol CaO}}{2 \text{ mol } SO_2} \times \frac{56.08 \text{ g CaO}}{1 \text{ mol CaO}} = 3.0 \times 10^3 \text{ g CaO}$$

Step 3: Think about your result.

The resultant mass could also be reported as 3.0 kg, with two significant figures. Even though the 2:2 mole ratio does not mathematically affect the problem, it is still necessary for unit conversion.

Practice Problem

5. Sodium azide (NaN_3) is a compound that is used in automobile air bags (**Figure ??**). A collision triggers its rapid decomposition into sodium and nitrogen gas, which is the gas that fills the air bag.
   (a) The decomposition of 1.00 g of NaN_3 produces what volume of N_2 at STP?
   (b) What mass of NaN_3 is required to produce 250. L of N_2 at STP?

Other Stoichiometry

Stoichiometric conversions all involve mole ratios between substances in a balanced chemical equation. Problems that involve mass and/or the volume of a gas are very common and practical. However, a third "arm" of the mole road map could also be part of a problemnumber of representative particles of a substance. Using the equation in Sample Problem 12.7, we could determine the number of formula units of aluminum sulfate produced when 25.0 g of Al reacts.

$$25.0 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{1 \text{ mol } Al_2(SO_4)_3}{2 \text{ mol Al}} \times \frac{6.02 \times 10^{23} \text{ form. units } Al_2(SO_4)_3}{1 \text{ mol } Al_2(SO_4)_3} = 2.79 \times 10^{23} \text{ form. units } Al_2(SO_4)_3$$

Problems could potentially arise involving any combination of mass, volume, and number of representative particles. However, since particles of this size cannot actually be counted, and stoichiometry is used most often for lab-based situations, problems involving the number of particles are seldom encountered in the real world.

Summary of Stoichiometry

The flowchart below (**Figure ??**) illustrates the types of stoichiometry problems that we have seen in this chapter and that you will most often need to solve. Conversion (b) is always present in any stoichiometry problem, while the use of the conversions represented by (a), (c), (d), and (e) depend on the specific type of problem. Conversion (f) is unique to gaseous volume-volume problems in which the pressure and temperature are held constant.

FIGURE 1.5

A municipal propane tank in Austin, TX. The combustion of propane gas produces carbon dioxide and water vapor.

To learn more about stoichiometric calculations, watch the video lecture at http://www.khanacademy.org/science/chemistry/chemical-reactions-stoichiometry/v/stoichiometry.

An example of how to solve a stoichiometry problem can be seen at http://www.khanacademy.org/science/physics/thermodynamics/v/stoichiometry-example-problem-1.

A second example of how to solve a stoichiometry problem can be seen at http://www.khanacademy.org/science/physics/thermodynamics/v/stoichiometry-example-problem-2.

Lesson Summary

- All stoichiometry problems involve the use of a mole ratio to convert between moles of a given substance and moles of an unknown substance.

- In an ideal stoichiometry problem, the mass of any reactant or product can be calculated from the mass of any other reactant or product if the balanced equation is known.
- Since the molar volume of any gas at STP is constant, gas volumes can also be used in stoichiometric calculations. This allows for multiple types of problems where amounts can be indicated by mass, volume, or number of moles.

Lesson Review Questions

Reviewing Concepts

1. How many conversion factors are involved in each of the following stoichiometry problems, where A and B are two components of a chemical reaction, and all volumes are for gases at STP?
   a. moles of A $\rightarrow$ mass of B
   b. volume of A $\rightarrow$ volume of B
   c. mass of B $\rightarrow$ mass of A
   d. moles of B $\rightarrow$ volume of A
2. Why does a mass-mass stoichiometry problem require three steps, while a volume-volume stoichiometry problem only requires one step?

Problems

3. The double-replacement reaction that occurs when aqueous solutions of calcium chloride and silver nitrate are combined produces a solution of calcium nitrate and a precipitate of silver chloride. $CaCl_2(aq) + 2AgNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2AgCl(s)$
   a. How many moles of calcium chloride are needed to react completely with 2.28 moles of silver nitrate?
   b. If 0.0623 moles of $CaCl_2$ reacts completely, how many grams of AgCl are produced?
   c. In order to produce exactly 1.50 g of AgCl, how many moles of each of the two reactants should be used?
4. Silicon dioxide reacts with carbon upon heating to produce silicon carbide (SiC) and carbon monoxide. $SiO_2(s) + 3C(s) \rightarrow SiC(s) + 2CO(g)$
   a. What mass of carbon is required to react completely with 15.70 g of SiO_2?
   b. When 152 g of SiO_2 reacts with excess carbon, what mass of SiC is produced?
   c. If 42.2 g of CO were produced by this reaction, what mass of carbon must have reacted?
5. Butane (C_4H_{10}) combusts according to the following reaction: $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$ Assume no change in temperature or pressure for the following questions.
   a. What volume of O_2 is needed to combust 425 mL of butane?
   b. What volume of butane must be combusted to produce 729 L of CO_2?
   c. When 6.20 L of butane is combusted, what volumes of CO_2 and H_2O are produced?
6. Dissolving calcium carbonate in hydrochloric acid produces aqueous calcium chloride, carbon dioxide, and water. Assume the reaction takes place at STP. $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$
   a. What volume of CO_2 is produced by the reaction of 9.58 g $CaCO_3$?
   b. If 67.1 L of CO_2 is produced by this reaction, what mass of HCl must have reacted?
   c. If 0.812 mol of HCl is used in this reaction, what mass of $CaCO_3$ would be consumed? What volume of CO_2 would be produced?
7. Nitroglycerin is an explosive compound that decomposes into multiple gaseous products according to the following reaction: $4C_3H_5(NO_3)_3(s) \rightarrow 12CO_2(g) + 10H_2O(g) + 6N_2(g) + O_2(g)$

a. What mass of nitrogen gas at STP is produced when 0.314 g of nitroglycerin decomposes?
b. What volume of carbon dioxide at STP is produced in a reaction that also produces 2.25 moles of O_2?
c. What is the total volume of all gases produced at STP by the full decomposition of 10.0 g of nitroglycerin?

8. Iron rusts to form iron(III) oxide according to the following equation: $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$
   a. If 46.2 g of Fe_2O_3 are produced by this reaction, how many Fe atoms reacted?
   b. What volume of O_2 at STP is needed to fully react with 8.39×10^{24} atoms of iron?
   c. The complete reaction of 0.916 mol Fe will produce how many formula units of Fe_2O_3?
9. Zinc reacts with hydrochloric acid according to the following equation: $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$ In a certain experiment, a 3.77 g sample of impure zinc is reacted with excess hydrochloric acid. If 1.09 L of H_2 gas is collected at STP, what percentage of the original sample was zinc? Assume that the impurities do not react with HCl.

Further Reading / Supplemental Links

- Stoichiometry, (http://www.wisc-online.com/objects/ViewObject.aspx?ID=GCH1504)
- The Chem Collective –Stoichiometry Tutorials –Reaction Stoichiometry, (http://www.chemcollective.org/stoich/reaction_stoi.php)
- Chemical Reaction Stoichiometry, (http://www.chemical-stoichiometry.net/)
- You can view a video lab of acid base stoichiometry at http://www.youtube.com/watch?v=GNNuzEZS8sw.
  - The document that goes with this lab is available at http://www.dlt.ncssm.edu/core/Chapter6-Stoichiometry/Chapter6-Labs/Stoi_Acid-Base_HCl-H2SO4_web_version.doc.
- Here is another online stoichiometry lab. It consists of two videos. Part 1 is at http://www.dlt.ncssm.edu/core/Chapter6-Stoichiometry/Chapter6-Labs/2_component_system_anal1-lg.htm, and part 2 is at http://www.dlt.ncssm.edu/core/Chapter6-Stoichiometry/Chapter6-Labs/2_component_system_anal2-lg.htm.
  - The lab document for this is found at http://www.dlt.ncssm.edu/core/Chapter6-Stoichiometry/Chapter6-Labs/Analysis_of_a_Two_Comp_Sys_web_ver.doc.
- View a microscale stoiciometry lab at http://www.youtube.com/watch?v=vA3o38sFNjY.
  - The document for this lab can be found at http://www.dlt.ncssm.edu/core/Chapter6-Stoichiometry/Chapter6-Labs/Microscale_stoi_web_ver.doc.
- View a stoichiometry experiment involving sodium bicarbonate and hydrochloric acid at http://www.youtube.com/watch?v=vjVrIFScsls.
  - The document for this lab is found at http://www.dlt.ncssm.edu/core/Chapter6-Stoichiometry/Chapter6-Labs/NaHCO3-HCl_lab_web_ver.doc.

Points to Consider

This lesson dealt with ideal stoichiometry, where 100% of the reactants were converted to products. In the real world, many chemical reactions do not proceed entirely in this way.

- How can we calculate the amount of products that could be formed in a reaction when two or more reactants are combined in a ratio other than the mole ratio from the balanced equation?
- How can we express the extent to which a set of reactants is converted to products if it is less than 100%?

1.3 Limiting Reactant and Percent Yield

Lesson Objectives

- Analyze a chemical reaction in order to determine which reactant is the limiting reactant and which is the excess reactant.
- Calculate the amount of excess reactant remaining after a reaction is complete.
- Calculate the theoretical yield of a reaction when the available amounts of each reactant are known.
- Calculate the percent yield of a reaction based on the theoretical and actual yields.

Lesson Vocabulary

- actual yield
- excess reactant (reagent)
- limiting reactant (reagent)
- percent yield
- theoretical yield

Check Your Understanding

Recalling Prior Knowledge

- What is the relationship between the amounts of reactants in a chemical reaction that follows ideal stoichiometry?
- How is the mass of a product determined when the mass of one reactant is known?

We have seen that the reactants in a chemical reaction combine with one another in a specific molar ratio that is described by the balanced equation. What if a reaction is performed in which the reactants are not present in that exact ratio? In this lesson, you will learn about this type of reaction and how to calculate the amounts of products that are expected to be formed.

Identifying the Limiting Reactant

Everyday Limiting Reactant

Cooking is a great example of everyday chemistry. In order to correctly follow a recipe, a cook needs to make sure that he has plenty of all the necessary ingredients in order to make his dish. Let us suppose that you are planning on making some pancakes for a large group of people (**Figure** 1.8).

FIGURE 1.8
When making pancakes, the number that you can make is limited by the amount of each ingredient that you have.

The recipe on the box indicates that the following ingredients are needed for each batch of pancakes:

- 1 cup of pancake mix
- $\frac{3}{4}$ cup milk
- 1 egg
- 1 tablespoon vegetable oil

Now you check the pantry and the refrigerator and see that you have the following ingredients available:

- 2 boxes of pancake mix (8 cups)
- Half gallon of milk (4 cups)
- 2 eggs
- Full bottle of vegetable oil (about 3 cups)

The question that you must ask is: How many batches of pancakes can I make? The answer is two. Even though you have enough pancake mix, milk, and oil to make many more batches of pancakes, you are *limited* by the fact that you only have two eggs. As soon as you have made two batches of pancakes, you will be out of eggs and your "reaction" will be complete.

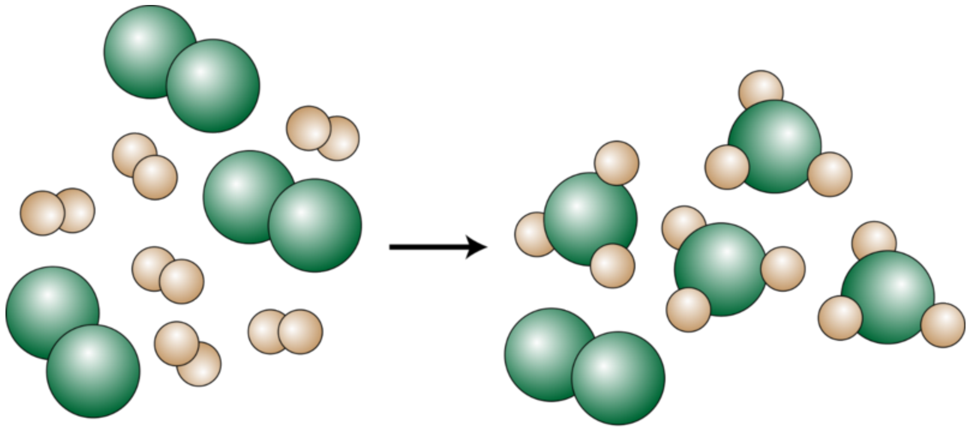
Limiting Reactants in Chemical Reactions

For a chemist, the balanced chemical equation is the recipe that must be followed. As you have seen earlier, the Haber process is a reaction in which nitrogen gas is combined with hydrogen gas to form ammonia. The balanced equation is shown below.

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

We know that the coefficients of the balanced equation tell us the mole ratio that is required for this reaction to occur. One mole of N_2 will react with three moles of H_2 to form two moles of NH_3.

Now let us suppose that a chemist were to react three moles of N_2 with six moles of H_2.



	Reactants		Products
Before reaction	3 mol N_2	6 mol H_2	0 mol NH_3
After reaction	1 mol N_2	0 mol H_2	4 mol NH_3

What happened in this reaction? The chemist started with 3 moles of N_2. You may think of this as being enough to make 3 batches of the "recipe" (as shown in the balanced equation), since the coefficient for N_2 is 1. However, the 6 moles of H_2 that the chemist started with is only enough for 2 batches, since the coefficient for H_2 is 3, and $3 \times 2 = 6$. After this reaction is complete, all of the hydrogen gas will be gone, but there will be 1 mole of nitrogen gas left over. At most, 4 moles of NH_3 (two "batches") can be generated with these amounts of reactants, because, after that point, one of the "ingredients" has been completely used up. The overall reaction that occurred is the following:

$$2 \text{ mol } N_2 + 6 \text{ mol } H_2 \rightarrow 4 \text{ mol } NH_3$$

All the amounts are doubled from the original balanced equation.

The **limiting reactant** (or **limiting reagent**) *is the reactant that determines the amount of product that can be formed in a chemical reaction*. The reaction proceeds until the limiting reactant is completely used up. In our example above, H_2 is the limiting reactant. The **excess reactant** (or **excess reagent**) *is any reactant that cannot be completely consumed by the full reaction of the limiting reactant*. In other words, there is always excess reactant left over after the reaction is complete. In the above example, N_2 is the excess reactant.

Get help with limiting reactant problems at http://www.chem.tamu.edu/class/majors/tutorialnotefiles/limiting.htm.

Watch an animation of a reaction involving a limiting reactant at http://www.dlt.ncssm.edu/core/Chapter6-Stoichiometry/Chapter6-Animations/LimitingReactant.html.

Solving Limiting Reactant Problems

In the real world, amounts of reactants and products are typically measured by mass or by volume. It is first necessary to convert the given quantities of each reactant to moles in order to identify the limiting reactant.

Sample Problem 12.9: Determining the Limiting Reactant

Silver metal reacts with sulfur to form silver sulfide according to the following balanced equation:

$$2Ag(s) + S(s) \rightarrow Ag_2S(s)$$

What is the limiting reactant when 50.0 g Ag is reacted with 10.0 g S?

Step 1: List the known quantities and plan the problem.

Known

- given: 50.0 g Ag
- given: 10.0 g S

Unknown

- limiting reactant

Use the atomic masses of Ag and S to determine the number of moles of each present. Then, use the balanced equation to calculate the number of moles of sulfur that would be needed to react with the number of moles of silver present. Compare this result to the actual number of moles of sulfur present.

Step 2: Solve.

First, calculate the number of moles of Ag and S present:

$$50.0 \text{ g Ag} \times \frac{1 \text{ mol Ag}}{107.87 \text{ g Ag}} = 0.464 \text{ mol Ag}$$
$$10.0 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = 0.312 \text{ mol S}$$

Second, find the moles of S that would be required to react with all of the given Ag:

$$0.464 \text{ mol Ag} \times \frac{1 \text{ mol S}}{2 \text{ mol Ag}} = 0.232 \text{ mol S (required)}$$

The amount of S actually present is 0.312 moles. The amount of S that is required to fully react with all of the Ag is 0.232 moles. Since there is more sulfur present than what is required to react, the sulfur is the excess reactant. Therefore, silver is the limiting reactant.

Step 3: Think about your result.

The balanced equation indicates that the necessary mole ratio of Ag to S is 2:1. Since there were not twice as many moles of Ag present in the original amounts, that makes silver the limiting reactant.

There is a very important point to consider about the preceding problem. Even though the mass of silver present in the reaction (50.0 g) was greater than the mass of sulfur (10.0 g), silver was the limiting reactant. This is because chemists must always convert to molar quantities and consider the mole ratio from the balanced chemical equation.

Practice Problems

1. The equation for the combustion of propane (C_3H_8) is: $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$. If 1.30 mol C_3H_8 is reacted with 6.00 mol O_2, identify the limiting reactant.
2. When zinc metal reacts with hydrochloric acid, zinc chloride and hydrogen gas are produced: $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$. Identify the limiting reactant when 8.00 g Zn reacts with 8.00 g HCl.

There are two other things that we would like to be able to determine in a limiting reactant problem. One is the quantity of the excess reactant that will be left over after the reaction is complete. Another is the quantity of product that will be formed in the reaction. We will go back to Sample Problem 12.9 to answer these questions.

Sample Problem 12.10A: Determining the Amount of Excess Reactant Left Over

What is the mass of excess reactant remaining when 50.0 g Ag reacts with 10.0 g S?

$$2Ag(s) + S(s) \rightarrow Ag_2S(s)$$

Step 1: List the known quantities and plan the problem.

Known

- Excess reactant = 0.312 mol S (from Sample Problem 12.9)
- Amount of excess reactant needed = 0.232 mol S (from Sample Problem 12.9)

Unknown

- Mass of excess reactant remaining after the reaction = ? g

Subtract the amount (in moles) of the excess reactant that will react from the amount that is originally present. Convert moles to grams.

Step 2: Solve.

$$0.312 \text{ mol S} - 0.232 \text{ mol S} = 0.080 \text{ mol S (remaining after reaction)}$$
$$0.080 \text{ mol S} \times \frac{32.07 \text{ g S}}{1 \text{ mol S}} = 2.57 \text{ g S}$$

There are 2.57 g of sulfur remaining when the reaction is complete.

Step 3: Think about your result.

There were 10.0 g of sulfur present before the reaction began. If 2.57 g of sulfur remains after the reaction, then 7.43 g S reacted.

$$7.43 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = 0.232 \text{ mol S}$$

This is the amount of sulfur that reacted. The problem is internally consistent.

Sample Problem 12.10B: Determining the Quantity of Product Formed in a Reaction

What mass of Ag_2S will be produced when 50.0 g Ag reacts with 10.0 g S?

$$2Ag(s) + S(s) \rightarrow Ag_2S(s)$$

Step 1: List the known quantities and plan the problem.

Known

- Limiting reactant = 0.464 mol Ag (from Sample Problem 12.9)
- Molar mass of Ag_2S = 247.81 g/mol

Unknown

- Mass of silver sulfide produced = ? g

The limiting reactant is the reactant that will determine the amount of product that is produced. Use stoichiometry to calculate the number of moles of Ag_2S produced and then convert that amount to grams.

$$\text{mol Ag} \rightarrow \text{mol Ag}_2\text{S} \rightarrow \text{g Ag}_2\text{S}$$

Step 2: Solve.

$$0.464 \text{ mol Ag} \times \frac{1 \text{ mol Ag}_2\text{S}}{2 \text{ mol Ag}} \times \frac{247.81 \text{ g Ag}_2\text{S}}{1 \text{ mol Ag}_2\text{S}} = 57.5 \text{ g Ag}_2\text{S}$$

There is 57.5 g Ag_2S formed in the reaction.

Step 3: Think about your result.

Silver is the limiting reactant, so all 50.0 grams of silver are used up in the reaction. Therefore, the mass of Ag_2S produced must be larger than 50.0 g. However, the mass of Ag_2S produced must be less than 60.0 g because the sulfur was in excess, so some of it did not react.

Practice Problems

3. Sodium chloride can be produced by reacting solid sodium with chlorine gas: In a certain reaction, 3.40 moles of Na are reacted with 4.20 moles of Cl_2.
   (a) Identify the limiting reactant.
   (b) Calculate the moles of excess reactant remaining after the reaction.
   (c) Calculate the moles of NaCl produced.
4. Using the Haber process for the production of ammonia: How many grams of ammonia will be produced when 35.0 g N_2 is reacted with 12.0 g H_2?

Watch a video lecture about limiting reactants at http://www.khanacademy.org/science/chemistry/chemical-reactions-stoichiometry/v/stoichiometry–limiting-reagent.

View another example of solving a limiting reactant problem at http://www.khanacademy.org/science/physics/thermodynamics/v/limiting-reactant-example-problem-1.

Percent Yield

Chemical reactions in the real world don't always go exactly as planned on paper. In the course of an experiment, many things will contribute to the formation of less product than would be predicted. Besides spills and other experimental errors, there are usually losses due to incomplete reactions, undesirable side reactions, and numerous other factors. Chemists need a measurement that indicates how successful a reaction has been. This measurement is called the percent yield.

To compute the percent yield, it is first necessary to determine how much of the product should be formed based on stoichiometry. This is called the **theoretical yield**, *the maximum amount of product that could be formed from the given amounts of reactants*. The **actual yield** *is the amount of product that is actually formed when the reaction is carried out in the laboratory*. The **percent yield** *is the ratio of the actual yield to the theoretical yield, expressed as a percentage*.

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

Percent yield is very important in the manufacture of products. Much time and money is spent improving the percent yield for chemical production. When complex chemicals are synthesized by many different reactions, one step with a low percent yield can quickly cause a large waste of reactants and money.

Percent yields are generally less than 100% because of the reasons indicated earlier. However, percent yields that appear to be greater than 100% are possible if the measured product of the reaction contains impurities that add to the mass of the pure product. When a chemist synthesizes a desired chemical, he or she must always be careful to purify the products of the reaction.

Sample Problem 12.11: Calculating the Theoretical Yield and the Percent Yield

Potassium chlorate decomposes when heated in the presence of a catalyst according to the reaction below:

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

In a certain experiment, 40.0 g $KClO_3$ is heated until it completely decomposes. What is the theoretical yield of oxygen gas? The experiment is performed, and the mass of the collected oxygen gas is found to be 14.9 g. What is the percent yield for the reaction?

Part A: First, we will calculate the theoretical yield based on the stoichiometry.

Step 1: List the known quantities and plan the problem.

Known

- given: mass of $KClO_3$ = 40.0 g
- molar mass of $KClO_3$ = 122.55 g/mol
- molar mass of O_2 = 32.00 g/mol

Unknown

- theoretical yield of O_2 = ? g

Apply stoichiometry to convert from the mass of a reactant to the mass of a product:

$$\text{g } KClO_3 \rightarrow \text{mol } KClO_3 \rightarrow \text{mol } O_2 \rightarrow \text{g } O_2$$

Step 2: Solve.

$$40.0 \text{ g } KClO_3 \times \frac{1 \text{ mol } KClO_3}{122.55 \text{ g } KClO_3} \times \frac{3 \text{ mol } O_2}{2 \text{ mol } KClO_3} \times \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} = 15.7 \text{ g } O_2$$

The theoretical yield of O_2 is 15.7 g.

Step 3: Think about your result.

The mass of oxygen gas must be less than the 40.0 g of potassium chlorate that was decomposed.

Part B: Now we use the actual yield and the theoretical yield to calculate the percent yield.

Step 1: List the known quantities and plan the problem.

Known

- Actual yield = 14.9 g
- Theoretical yield = 15.7 g (from ***Part A***)

Unknown

- Percent yield = ? %

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

Use the percent yield equation above.

Step 2: Solve.

$$\text{Percent Yield} = \frac{14.9 \text{ g}}{15.7 \text{ g}} \times 100\% = 94.9\%$$

Step 3: Think about your result.

Since the actual yield is slightly less than the theoretical yield, the percent yield is just under 100%.

Practice Problems

5. If 23.5 g of lithium nitride is reacted with an excess of water, 40.2 g of lithium hydroxide is produced according to the following reaction: What is the percent yield of the reaction?
6. Under a certain set of conditions, the Haber process is known to proceed such that the percent yield for the reaction is 86.0%. If the actual yield is 27.8 g NH_3, what is the theoretical yield for the reaction?

Practice more percent yield problems at http://www.uen.org/utahlink/tours/tourElement.cgi?element_id=42288&tour_id=17891&category_id=33176.

You can find more practice problems at http://science.widener.edu/svb/tutorial/percentyieldcsn7.html.

Lesson Summary

- The limiting reactant in a chemical reaction is the reactant that determines the amount of product that can be formed. More of the excess reactant is present than is needed to completely react with the limiting reactant, so some of the excess reactant remains after the reaction is complete.
- The theoretical yield is the maximum amount of product that could be formed based on stoichiometry calculations.
- The percent yield is a measure of the efficiency of the reaction and is calculated by dividing the actual yield by the theoretical yield and converting to a percentage.

Lesson Review Questions

Reviewing Concepts

1. What is wrong with this statement? "The limiting reactant of a reaction is whichever reactant is present in a smaller amount."
2. Which reactant, limiting or excess, determines the theoretical yield of products?
3. Why are actual yields generally less than theoretical yields?
4. What is likely to have occurred if a calculated percent yield is larger than 100%?

Problems

5. In a certain reaction, 0.740 mol of lithium is combined with 0.180 mol of nitrogen gas. The reaction forms lithium nitride according to the following equation: $6Li(s) + N_2(g) \rightarrow 2Li_3N(s)$
   a. Identify the limiting reactant.
   b. Calculate the number of moles of excess reactant remaining after the reaction is complete.
   c. Calculate the number of moles of product formed.
6. The reaction of fluorine gas with ammonia produces dinitrogen tetrafluoride and hydrogen fluoride. $5F_2(g) + 2NH_3(g) \rightarrow N_2F_4(g) + 6HF(g)$ In a certain reaction, 45.8 g F_2 is reacted with 20.6 g NH_3.

a. Determine the limiting reactant.
b. How many grams of the excess reactant remains after the reaction is complete?
c. What is the theoretical yield of each of the two products?
d. The reaction is performed and 16.2 g of N_2F_4 is produced. Calculate the percent yield for the reaction.

7. 26.3 g of potassium metal reacts with 4.89 L of oxygen gas at STP, forming potassium oxide. $4K(s) + O_2(g) \rightarrow 2K_2O(s)$

a. Determine the limiting reactant.
b. How much of the excess reactant remains after the reaction is complete? If the excess reactant is oxygen, give its volume.
c. What mass of potassium oxide could theoretically be produced?

8. The percentage yield for the reaction of PCl_3 with Cl_2 to form PCl_5 is 80.8%. What mass of PCl_5 would be produced by the reaction of 37.8 g of PCl_3 with excess Cl_2? $PCl_3(g) + Cl_2(g) \rightarrow PCl_5(g)$
9. If 26.7 g of nickel metal is reacted with 178 g of silver nitrate, what mass of silver metal will be produced? What is the limiting reactant? $Ni(s) + 2AgNO_3(aq) \rightarrow Ni(NO_3)_2(aq) + 2Ag(s)$
10. The Ostwald Process is a series of three reactions (shown below) that produces nitric acid from ammonia. 1.00 kg of ammonia is subjected to the reactions, and unlimited amounts of oxygen and water are available. If each reaction proceeds at a 92.5% yield, what mass of HNO_3 is produced? Assume that any NO produced by the third reaction is lost and not reused as a reactant in the second reaction.

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$$
$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$
$$3NO_2(g) + H_2O(g) \rightarrow 2HNO_3(aq) + NO(g)$$

Further Reading / Supplemental Links

- The Limiting Reagent in Chemical Reactions, (http://www.wisc-online.com/objects/ViewObject.aspx?ID=GCH7404)
- Product Yields in Chemical Reactions, (http://www.wisc-online.com/objects/ViewObject.aspx?ID=GCH7504)
- The Chem Collective –Stoichiometry Tutorials –Limiting Reagents, (http://www.chemcollective.org/stoich/limiting-reagents.php)
- The Chem Collective –Stoichiometry Tutorials –Product Formation, (http://www.chemcollective.org/stoich/percentyield.php)
- You can use stoichiometry to determine empirical and molecular formula of an unidentified reactant. This combines skills from this section and the section on empirical and molecular formulas. Watch an example at http://www.khanacademy.org/science/physics/thermodynamics/v/empirical-and-molecular-formulas-from-stoichiometry.
- Another example of using stoichiometry to determine the empirical and molecular formula of an unidentified reactant is at http://www.khanacademy.org/science/physics/thermodynamics/v/example-of-finding-reactant-empirical-formula.

Points to Consider

Now that you have an understanding of the quantitative relationships that exist between chemical substances in chemical reactions, we will turn our attention to the states of matter and their relationships to one another.

- What are the three states of matter?
- How can changes in temperature and pressure affect the state of a sample of matter?

1.4 References

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