Section 12.3

Limiting Reactants

Objectives

- **Identify** the limiting reactant in a chemical equation.
- **Identify** the excess reactant and **calculate** the amount remaining after the reaction is complete.
- **Calculate** the mass of a product when the amounts of more than one reactant are given.

Vocabulary

limiting reactant
excess reactant

At a school dance, the music begins and boys and girls pair up to dance. If there are more boys than girls, some boys will be left without partners. The same is true of reactants in a chemical reaction. Rarely in nature are reactants in a chemical reaction present in the exact ratios specified by the balanced equation. Generally, one or more reactants are in excess and the reaction proceeds until all of one reactant is used up.

Why do reactions stop?

When a chemical reaction is carried out in the laboratory, the same principle applies. Usually, one or more reactants are in excess, while one is limited. The amount of product depends upon the reactant that is limited.

Remember the reaction between potassium permanganate and sodium hydrogen sulfite in the **DISCOVERY LAB.** As you added colorless sodium hydrogen sulfite to purple potassium permanganate, the color faded as a reaction took place. Finally, the solution was colorless. You could have continued adding sodium hydrogen sulfite, but would any further reaction have taken place? You are correct if you said that no further reaction could take place because no potassium permanganate was available to react. Potassium permanganate was a limiting reactant. As the name implies, the **limiting reactant** limits the extent of the reaction and, thereby, determines the amount of product. A portion of all of the other reactants remains after the reaction stops. These left-over reactants are called **excess reactants**. What was the excess reactant in the reaction of potassium permanganate and sodium hydrogen sulfite?

To help you understand limiting reactants, consider the analogy in **Figure 12-5**. How many tool sets can be assembled from the items shown if each complete tool set consists of one pair of pliers, one hammer, and two screwdrivers? You can see that four complete tool sets can be assembled. The number of tool sets is limited by the number of available hammers. Pliers and screwdrivers remain in excess. Chemical reactions work in a similar way.

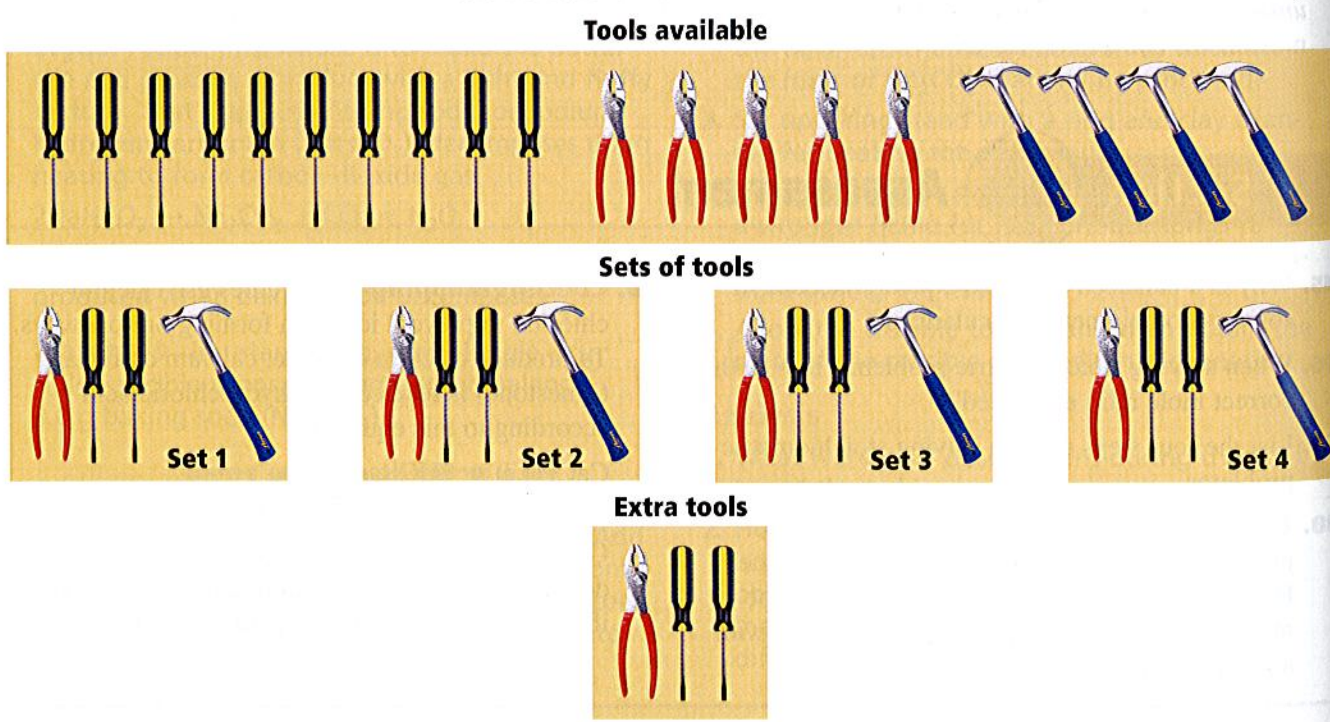


Figure 12-5

Each tool set must have one hammer so only four sets can be assembled. Which tool is limiting? Which tools are in excess?

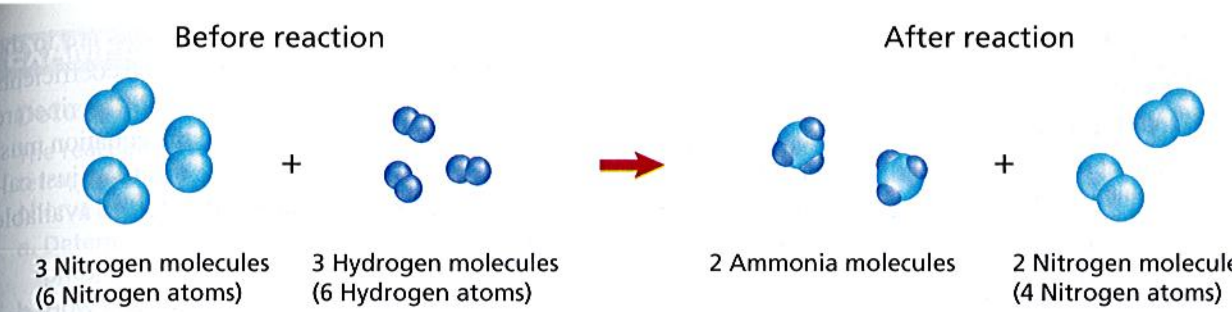


Figure 12-6

Refer to **Table C-1** in Appendix C for a key to atom color conventions. Check to see whether all the atoms present before the reaction are present after the reaction. Some nitrogen molecules were unchanged in the reaction. Which reactant is in excess?

The calculations you did in Section 12.2 were based on having the reactants present in the ratio described by the balanced chemical equation. How can you calculate the amount of product formed when one reactant limits the amount of product and the other is in excess? The first thing you must do is determine which reactant is the limiting reactant.

Consider the reaction shown in **Figure 12-6** in which three molecules of nitrogen (N_2) and three molecules of hydrogen (H_2) react to form ammonia (NH_3). You can visualize that in the first step of the reaction, all the nitrogen molecules and hydrogen molecules are separated into individual atoms. These are the atoms available for reassembling into ammonia molecules just like the tools in **Figure 12-5** before they were assembled into tool kits. How many molecules of ammonia will be produced from the available atoms? Four tool kits could be assembled from the tools because only four hammers were available. Two ammonia molecules can be assembled from the hydrogen and nitrogen atoms because only six hydrogen atoms are available, three for each ammonia molecule. When the hydrogen is gone, two molecules of nitrogen remain unreacted. Thus, hydrogen is the limiting reactant and nitrogen is the excess reactant. It's important to know which reactant is the limiting reactant because, as you have just learned, the amount of product formed depends upon this reactant.

ChemistryOnline

Topic: Limiting Reactants
To learn more about limiting reactants, visit the Chemistry Web site at **chemistrymc.com**

Activity: Research and explain how nutrients in soil can be depleted when crops are grown on a plot of land year after year. What is the limiting reactant in this case? How may the soil be enriched?

Percent Yield

Objectives

- **Calculate** the theoretical yield of a chemical reaction from data.
- **Determine** the percent yield for a chemical reaction.

Vocabulary

theoretical yield
actual yield
percent yield

Suppose you were determined to improve your jump shot and took time each afternoon to practice. One afternoon, you succeeded in getting the ball through the hoop 49 times out of a total of 75 tries. Theoretically, you could have been successful 75 times, but in actuality that usually doesn't happen. (That's why you practice.) But how successful were you? You could calculate your efficiency as a percent by dividing the number of successful tries by the total number of tries and multiplying by 100.

$$\frac{49 \text{ actual successes}}{75 \text{ theoretical successes}} \times 100 = 65\% \text{ successful shots}$$

Sixty-five percent successful jump shots means that you could expect to get the ball into the basket 65 times if you made 100 attempts.

Similar calculations are made to determine the success of chemical reactions because most reactions never succeed in producing the predicted amount of product. Although your work with stoichiometric problems so far may have led you to think that chemical reactions proceed according to the balanced equation without any difficulties and always produce the calculated amount of product, this is not the case! Not every reaction goes cleanly or completely. Many reactions stop before all of the reactants are used up, so the actual amount of product is less than expected. Liquid reactants or products may adhere to the surfaces of containers or evaporate, and solid product is always left behind on filter paper or lost in the purification process. In some instances, products other than the intended ones may be formed by competing reactions, thus reducing the yield of the desired product.

Math Handbook

Review the meaning and calculation of percents in the **Math Handbook** on page 907 in this text.

How much product?

In many of the calculations you have been practicing, you have been asked to calculate the amount of product that can be produced from a given amount of reactant. The answer you obtained is called the theoretical yield of the reaction. The **theoretical yield** is the maximum amount of product that can be produced from a given amount of reactant. A chemical reaction rarely produces the theoretical yield of product. A chemist determines the actual yield of a reaction through a careful experiment in which the mass of the product is measured. The **actual yield** is the amount of product actually produced when the chemical reaction is carried out in an experiment.

Chemists need to know how efficient a reaction is in producing the desired product. One way of measuring efficiency is by means of percent yield. Just as you calculated your percent of successful jump shots, a chemist can calculate what percent of the amount of product that could theoretically be produced was actually produced. **Percent yield** of product is the ratio of the actual yield to the theoretical yield expressed as a percent.

$$\text{Percent yield} = \frac{\text{actual yield (from an experiment)}}{\text{theoretical yield (from stoichiometric calculations)}} \times 100$$

The **problem-solving LAB** on page 372 will help you understand the importance of percent yield in chemical reactions and the kind of factors that may determine the size of the percent yield.

EXAMPLE PROBLEM 12-6

Calculating Percent Yield

When potassium chromate (K_2CrO_4) is added to a solution containing 0.500 g silver nitrate ($AgNO_3$), solid silver chromate (Ag_2CrO_4) is formed.

a. Determine the theoretical yield of the silver chromate precipitate.

b. If 0.455 g of silver chromate is obtained, calculate the percent yield.

1. Analyze the Problem $2AgNO_3 + K_2CrO_4 \rightarrow Ag_2CrO_4 + 2KNO_3$

You are given the mass of the reactant $AgNO_3$ and the actual yield of the product Ag_2CrO_4. You need to write the balanced chemical equation and calculate the theoretical yield by making these conversions: grams of silver nitrate to moles of silver nitrate, moles of silver nitrate to moles of silver chromate, moles of silver chromate to grams of silver chromate. The percent yield can be calculated from the actual yield of product and the calculated theoretical yield.

Known

mass of silver nitrate = 0.500 g $AgNO_3$
actual yield = 0.455 g Ag_2CrO_4

Unknown

theoretical yield = ? g Ag_2CrO_4
percent yield = ? % Ag_2CrO_4

2. Solve for the Unknown

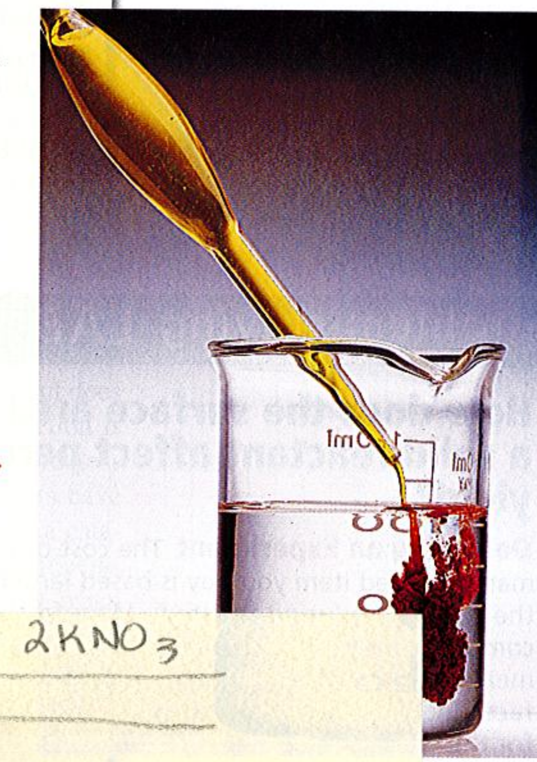
Write the balanced chemical equation and indicate the known and unknown quantities.

$$\overset{0.500\text{ g}}{2AgNO_3(aq)} + K_2CrO_4(aq) \rightarrow \overset{?\text{ g}}{Ag_2CrO_4(s)} + 2KNO_3(aq)$$

	$2AgNO_3$	K_2CrO_4	$\rightarrow Ag_2CrO_4$	$2KNO_3$
g	0.500		0.488 *	
g/mol	169.9		331.7	
mol	2.94×10^{-3}		1.47×10^{-3}	

$(\frac{1}{2})$

$$\% \text{ yield} = \frac{0.455\text{ g}}{0.488\text{ g}} \times 100 = 93.2\%$$



...ll ...s are ...er ...in ...mate ...potassium ...orange)

3. Evaluate the Answer

All quantities have three significant figures so the percent is correctly stated with three digits. The molar mass of Ag_2CrO_4 is about twice the molar mass of $AgNO_3$, and the ratio of mol $AgNO_3$ to mol Ag_2CrO_4 in the equation is 2:1. Therefore, 0.500 g $AgNO_3$ should produce about the same mass of Ag_2CrO_4. The actual yield of Ag_2CrO_4 is close to 0.500 g, so a percent yield of 93.2% is reasonable.

Practice!

For more practice calculating percent yield, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

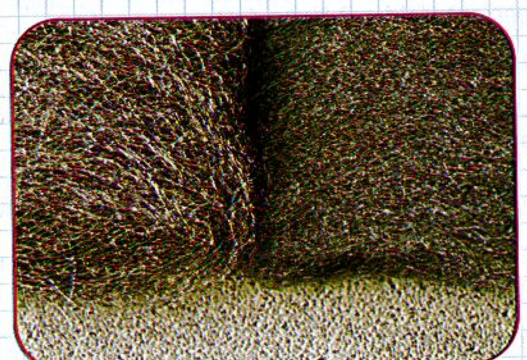
27. Aluminum hydroxide is often present in antacids to neutralize stomach acid (HCl). If 14.0 g aluminum hydroxide is present in an antacid tablet, determine the theoretical yield of aluminum chloride produced when the tablet reacts with stomach acid. If the actual yield of aluminum chloride from this tablet is 22.0 g, what is the percent yield?

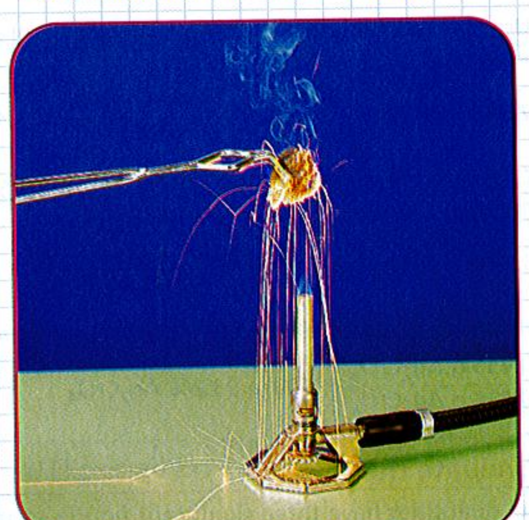
$Al(OH)_3(s) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)$

problem-solving LAB

How does the surface area of a solid reactant affect percent yield?

Designing an Experiment The cost of every manufactured item you buy is based largely on the cost of producing the item. Manufacturers compete to reduce costs and increase profits. This means increasing the percent yield of the manufacturing process by producing the most product for the amount of reactant used. If you were going to produce iron oxide (Fe_2O_3) from steel wool, how could you design an experiment to determine what gauge (diameter) steel wool will produce the highest yield? Write the equation for the reaction upon which you will base your experiment.





Analysis

In the first photo, different gauges of steel wool are shown. The second photo shows the combustion of a sample of steel wool using a Bunsen burner.

Thinking Critically

1. What quantities must be used to calculate the percent yield of Fe_2O_3 when iron is burned? How will these quantities be measured and how many measurements should be made?
2. What quantities should be kept constant in the experiment?
3. How will the resulting data be analyzed?
4. Are there any obvious errors in the design that could significantly affect the results? If so, how could they be avoided?

Percent yield in the marketplace You learned in the **problem-solving LAB** that in order to compete, manufacturers must reduce the cost of making their products to the lowest level possible. Percent yield is important in the calculation of overall cost effectiveness in industrial processes. For example, sulfuric acid (H_2SO_4) is made using mined sulfur, **Figure 12-8.** Sulfuric acid is an important chemical because it is a raw material for products such as fertilizers, detergents, pigments, and textiles. The cost of sulfuric acid affects the cost of many of the consumer items you use every day.

A two-step process called the contact process is often used for the manufacture of sulfuric acid. Over time, the process has been improved by chemical engineers to produce the maximum yield of product and, at the same time, comply with environmental standards for clean air. The two steps in the contact process are

$$S_8(s) + 8O_2(g) \rightarrow 8SO_2(g)$$

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

A final step, the combination of SO_3 with water, produces the product, H_2SO_4.

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

The first step, the combustion of sulfur, produces almost 100% yield. The second step also produces a high yield if a catalyst is used at the relatively low temperature of 400°C. A catalyst is a substance that speeds a reaction but does not appear in the chemical equation. Under these conditions, the reaction is slow. Raising the temperature speeds up the reaction but the yield decreases.

To maximize yield and minimize time in the second step, engineers have devised a system in which the reactants, O_2 and SO_2, are passed over a catalyst at 400°C. Because the reaction releases a great deal of heat, the temperature gradually increases with an accompanying decrease in yield. Thus, when the temperature reaches approximately 600°C, the mixture is cooled and then passed over the catalyst again. A total of four passes over the catalyst with cooling between passes results in a yield greater than 98%. This four-pass procedure maximizes the yield at temperatures near 400°C, and uses the modest increase in temperature to increase the rate and minimize the time.

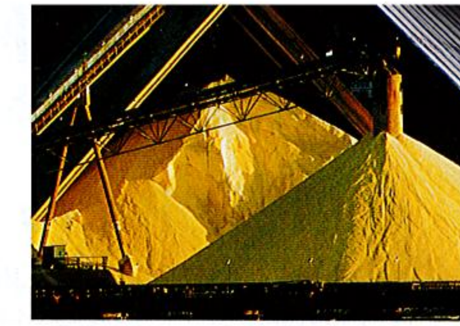


Figure 12-8

The large tonnage of sulfur needed to make sulfuric acid and other products is often obtained by forcing hot water into underground deposits to melt the sulfur and then pumping the liquid sulfur to the surface.

Section 12.4 Assessment

30. Distinguish between the theoretical yield and the actual yield of a chemical reaction.

31. Give several reasons why the actual yield is not usually equal to the theoretical yield.

32. Explain how percent yield is calculated.

33. **Thinking Critically** In an experiment, you are to combine iron with an excess of sulfur and heat the mixture to obtain iron(III) sulfide.

$2Fe(s) + 3S(s) \rightarrow Fe_2S_3(s)$

What experimental information must you collect in order to calculate the percent yield of this reaction?

34. **Interpreting Data** Use the data to determine the percent yield of the following reaction.

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

Oxygen is in excess.

Reaction Data	
Mass of crucible	35.67 g
Mass of crucible + Mg	38.06 g
Mass of Mg	
Mass of crucible + MgO	39.15 g
Mass of MgO	

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