7.3 Percentage Yield

Section Preview/ Specific Expectations

In this section, you will

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

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

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

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

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

Theoretical Yield and Actual Yield

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

CHEM FACT

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

Why Actual Yield and Theoretical Yield Are Often Different

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

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

Here are the chemical equations for these competing reactions:

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

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

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

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

Calculating Percentage Yield

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

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

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

Sample Problem

Calculating Percentage Yield

Problem

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

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

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

What Is Required?

You need to find the percentage yield of the reaction.

What Is Given?

actual yield = 1.72 g
theoretical yield = 9.10 g

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Plan Your Strategy

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

Act on Your Strategy

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

The percentage yield of the reaction is 18.9%.

Check Your Solution

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

Practice Problems

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

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

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

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

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

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

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

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

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

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

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

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

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

Sample Problem

Predicting Actual Yield Based on Percentage Yield

Problem

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

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

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

What Is Required?

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

What Is Given?

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

Plan Your Strategy

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

Act on Your Strategy

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

0.124 mol → 0.124 mol

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

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

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

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

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

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

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

12.4 g $CaCO_3$ → 6.95 g CaO

Check Your Solution

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

Continued ...

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Practice Problems

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

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

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

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

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

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

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

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

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

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

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

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

mind STRETCH

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

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

COURSE CHALLENGE

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

Applications of Percentage Yield

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

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

Careers in Chemistry

Chemical Engineer

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

Producing More for Less

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

Becoming a Chemical Engineer

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

Make Career Connections

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

Percentage Purity

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

SKILL FOCUS

Predicting

Performing and recording

Analyzing and interpreting

Investigation 7-B

Determining the Percentage Yield of a Chemical Reaction

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

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

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

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

Question

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

Predictions

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

Materials

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

Safety Precautions

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

Procedure

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

Observations

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

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

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

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

Analysis

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

Conclusion

3. Calculate the percentage yield for this reaction.

Applications

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

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

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

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

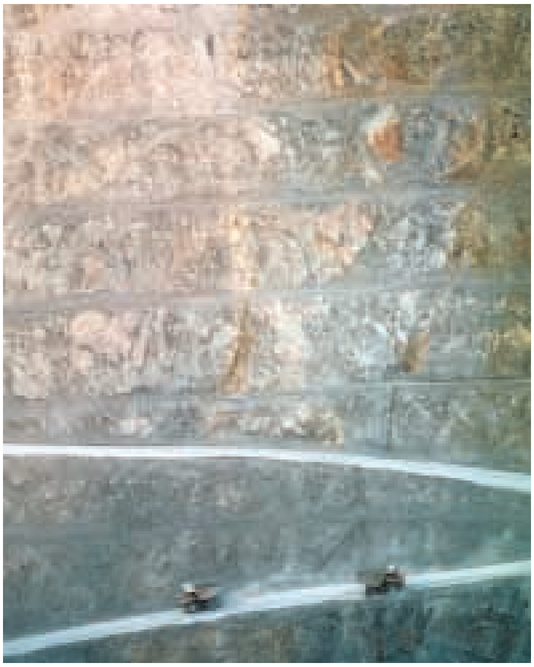

Figure 7.9 Copper is removed from mines like this one in the form of an ore. There must be sufficient copper in the ore to make the mine economically viable.

Sample Problem

Finding Percentage Purity

Problem

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

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

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

What Is Required?

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

What Is Given?

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

Plan Your Strategy

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

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

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

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Act on Your Strategy

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

0.100 mol ⟵ 0.0502 mol

3 unknown ratio / known ratio

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

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

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

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

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

12.0 g FeS_2 — 8.02 g Fe_2O_3

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

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

Check Your Solution

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

Practice Problems

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

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

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

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

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

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

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

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

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

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

Section Wrap-up

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

Section Review

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

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

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

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

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

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

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

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