SECTION 7.3

Reaction Yields

Key Terms

theoretical yield
actual yield
competing reaction
percentage yield

The highest grade that a student can earn on a test is 100 percent. However, most students do not receive a grade of 100 percent. The actual grade earned is usually less than the theoretical highest grade of 100 percent. The actual percentage grade on a test is calculated using the following equation:

$$\text{percentage grade} = \frac{\text{points earned}}{\text{maximum possible points}} \times 100\%$$

Calculating your test grade is similar to calculating chemical reaction yields.

Theoretical Yield and Actual Yield

theoretical yield the amount of product that is predicted by stoichiometric calculations

actual yield the actual amount of product that is recovered after a reaction is complete

competing reaction a reaction that occurs along with the principal reaction and that involves the reactants and/or products of the principal reaction

The stoichiometric calculations you have learned so far in this chapter allow you to calculate the amount of product that forms, in theory, in a chemical reaction. This is called the **theoretical yield** of the reaction. However, the theoretical yield is not always the same as the amount of product that actually forms in a chemical reaction. Just as most students are unlikely to answer 100 out of 100 questions on a test correctly, most reactions do not produce the theoretical amount of product. The actual amount of product that forms in a chemical reaction conducted in a laboratory or in industry is called the **actual yield**.

Competing Reactions (Reactants) That Affect Yield

Reactions do not yield as much product as expected for a variety of reasons. For example, the actual yield might be affected by a competing reaction. A **competing reaction** is a reaction that occurs along with the principal reaction and that involves the reactants and/or products of the principal reaction. An example of a competing reaction occurs when hydrocarbons are burned as fuel. Depending on the availability of oxygen and other circumstances, a hydrocarbon gas, such as propane, may burn in both a complete combustion reaction (the principal reaction) and an incomplete combustion reaction (the competing reaction) according to the following chemical equations:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g) \text{ (complete combustion)}$$

$$2C_3H_8(g) + 7O_2(g) \rightarrow 6CO(g) + 8H_2O(g) \text{ (incomplete combustion)}$$

Because the competing reaction uses the same reactants as the principal reaction, the principal reaction has a lower yield.

One product of the incomplete combustion reaction above is carbon monoxide, $CO(g)$, an odourless toxic gas. A carbon monoxide detector, like the one shown in Figure 7.9, alerts people if the gas is present in their home.

Figure 7.9 Propane can produce carbon monoxide during incomplete combustion. Because carbon monoxide is poisonous, detectors are used in homes to warn people if the gas is present.

Competing Reactions (Products) That Affect Yield

A competing reaction does not always involve the reactants in the principal reaction. It can also involve the products. For example, phosphorus reacts with chlorine and forms phosphorus trichloride, as shown in **Figure 7.10**. Some of the phosphorus trichloride then reacts with chlorine and forms phosphorus pentachloride, resulting in less actual yield. The chemical equations for these reactions are

$$2P(s) + 3Cl_2(g) \rightarrow 2PCl_3(\ell)$$
$$PCl_3(\ell) + Cl_2(g) \rightarrow PCl_5(s)$$

In the first reaction, phosphorus and chlorine gas produce phosphorus trichloride. In the second reaction, some of the phosphorus trichloride is converted into phosphorus pentachloride. The second reaction is the competing reaction. It causes the actual yield of phosphorus trichloride to be less than the theoretical yield.

Figure 7.10 When phosphorus reacts with chlorine gas, a competing reaction lowers the actual yield of phosphorus trichloride.

Other Factors That Affect Yield

Competing reactions are not the only things that can affect the actual product yield. Reaction rates can also affect product yield. For example, if a reaction is extremely slow, the reaction might not have gone to completion at the time the product yield is measured. Because the reaction is incomplete, the product yield will be lower than predicted.

Reaction rates can be affected by factors such as surface area, temperature, pressure, and reaction vessel conditions. Typically, larger surface areas of the reactants and higher reaction temperatures and pressures result in more frequent collisions by the particles that make up the reactants. Increasing the collision rate of the particles of the reactants increases the rate at which product forms. A reaction vessel can hamper the collisions of particles of the reactants because of factors such as a rough, pitted, or dirty surface, or an odd shape. As a result, fewer products will be formed. Another factor that reduces actual product yield is the purity of the reactants. If theoretical yield calculations are made on the assumption that the reactants are 100 percent pure when, in fact, they are not, the actual yield will be lower than expected. For example, if you use household vinegar instead of pure acetic acid in a reaction, your actual product yield will be much lower than expected because household vinegar is only about 5 percent acetic acid. The remaining 95 percent is water.

The laboratory techniques used to collect the final product also affect product yield. For example, if a product is slightly soluble in a filtrate, some of the product will remain dissolved in the filtrate. Furthermore, if the product is soluble in water, some of the product might dissolve in the water used to clean the filter paper. Other laboratory techniques can also cause product loss, such as the product clinging to the glassware, stirring rods, and other equipment used in the investigation.

Suggested **Investigation**

Inquiry Investigation 7-B, Finding the Percentage Yield of a Single Displacement Reaction

Summary of Factors That Affect Product Yield Summary

There are many factors that can affect the actual product yield. **Table 7.1** summarizes several of those factors. When you perform the Investigations 7-B and 7-C, you will observe some of these factors.

Table 7.1 Factors That Affect the Actual Yield of a Reaction

Factor	Description
Competing reaction	• Competing reactions involving the principal reactants and/or products produce multiple products and reduce the yield of the desired product.
Reaction rate	• A slow reaction does not go to completion and reaction products are collected too soon. • The surface area of the reactants is small, which reduces the probability of collisions of the particles of the reactants. • Environmental conditions of the reaction, such as temperature and pressure, slow particle movement and reduce the probability of reactant particle collisions. • Reaction vessel conditions, such as rough, pitted, or dirty surfaces or an odd-shaped vessel, impede reactant particle collisions.
Purity of the reactant	• Impure reactants contain contaminants. • Incorrect theoretical yield calculations are based on impure reactants and the mass of the impurity is not considered.
Laboratory techniques	• Improper lab techniques reduce actual product yield. • A slightly soluble product results in some of the product staying in the filtrate rather than being collected on the filter paper. • A slightly soluble product dissolves in the water used to rinse the product on the filter paper; some of the product is washed away during the rinsing process. • Some of the product clings to the reaction vessel, filter paper, stirring rods, spatulas, and other equipment used in the investigation and this mass is not recorded.

Learning Check

13. Explain the difference between actual yield and theoretical yield.

14. Which is usually higher, the actual yield or the theoretical yield? Explain your answer.

15. Explain how laboratory techniques may reduce the actual yield.

16. Draw and label a diagram showing how a reaction vessel can influence the actual yield.

17. Explain how impure reactants can affect the actual yield.

18. Use an analogy to explain how a competing reaction can reduce the actual yield.

Calculating Percentage Yield

percentage yield the actual yield of a reaction, expressed as a percentage of the theoretical yield

Chemists often need to know the efficiency of a chemical reaction, such as when they are doing a cost analysis for a chemical process or when they are ordering materials for a chemical process. The **percentage yield** of a chemical reaction is the actual yield expressed as a percentage of the theoretical yield:

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

In this equation, the theoretical yield is the amount of product that is determined by theoretical stoichiometric calculations, and the actual yield is the amount of product that is actually recovered in the reaction. The percentage yield can be calculated using either the amount in moles or mass. The following Sample Problems demonstrate how to use this equation.

Sample Problem

Calculating Percentage Yield

Problem

The following chemical equation represents the production of ethanol, $C_2H_5OH(aq)$, by the fermentation of glucose, $C_6H_{12}O_6(aq)$:

$$C_6H_{12}O_6(aq) \longrightarrow 2C_2H_5OH(aq) + 2CO_2(g)$$

If 20.0 g of glucose reacts but only 1.40 g of ethanol is produced, what is the percentage yield of the reaction?

glucose

What Is Required?

You need to find the percentage yield of the reaction.

What Is Given?

You know the balanced chemical equation:

$C_6H_{12}O_6(aq) \longrightarrow 2C_2H_5OH(aq) + 2CO_2(g)$

You know the actual yield of ethanol: 1.40 g

You know the mass of glucose: 20.0 g

Plan Your Strategy	Act on Your Strategy
Find the molar masses of glucose and ethanol.	$M_{C_6H_{12}O_6} = 6M_C + 12M_H + 6M_O$ $= 6(12.01\text{ g/mol}) + 12(1.01\text{ g/mol}) + 6(16.00\text{ g/mol}) = 180.18\text{ g/mol}$ $M_{C_2H_5OH} = 2M_C + 6M_H + 1M_O$ $= 2(12.01\text{ g/mol}) + 6(1.01\text{ g/mol}) + 1(16.00\text{ g/mol}) = 46.08\text{ g/mol}$
Determine the actual yield amount in moles of ethanol.	$n = \frac{m}{M} = \frac{1.40\ \cancel{\text{g}}}{46.08\ \cancel{\text{g}}/\text{mol}} = 0.03038\text{ mol }C_2H_5OH$
To determine the theoretical yield, convert the mass of glucose in grams to amount in moles. Then, use the mole ratio to find the yield of ethanol in moles.	$n_{C_6H_{12}O_6} = 20.0\ \cancel{\text{g}}\left(\frac{1\text{ mol }C_6H_{12}O_6}{180.18\ \cancel{\text{g}}\ C_6H_{12}O_6}\right) = 0.111\text{ mol }C_6H_{12}O_6$ $n_{C_2H_5OH} = 0.111\ \cancel{\text{mol }C_6H_{12}O_6}\left(\frac{2\text{ mol }C_2H_5OH}{1\ \cancel{\text{mol }C_6H_{12}O_6}}\right) = 0.222\text{ mol }C_2H_5OH$
Calculate the percentage yield of ethanol.	$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{0.03038\ \cancel{\text{mol}}}{0.222\ \cancel{\text{mol}}} \times 100\%$ $= 13.7\%$ Therefore, the percentage yield is 13.7%.

Check Your Solution

To check your answer, calculate the percentage yield using mass and the factor-label method.

$$\text{theoretical yield}_{C_2H_5OH} = (20.0\ \cancel{\text{g }C_6H_{12}O_6}) \times \left(\frac{1\ \cancel{\text{mol }C_6H_{12}O_6}}{180.18\ \cancel{\text{g }C_6H_{12}O_6}}\right) \times \left(\frac{2\ \cancel{\text{mol }C_2H_5OH}}{1\ \cancel{\text{mol }C_6H_{12}O_6}}\right) \times \left(\frac{46.08\text{ g }C_2H_5OH}{1\ \cancel{\text{mol }C_2H_5OH}}\right)$$

$$= 10.2\text{ g }C_2H_5OH$$

$$\text{actual percentage yield}_{C_2H_5OH} = \frac{1.40\ \cancel{\text{g}}}{10.2\ \cancel{\text{g}}} \times 100\% = 13.7\%$$

The two methods produce the same answer.

Sample Problem

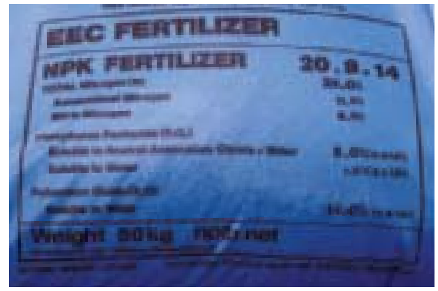
Predicting the Actual Yield Using Percentage Yield

Problem

Ammonium nitrate, $NH_4NO_3(s)$, is a compound that is used to make fertilizer, like the fertilizer shown on the right. It is produced in a chemical reaction that is represented by the following balanced equation:

$$NH_3(g) + HNO_3(aq) \rightarrow NH_4NO_3(s)$$

Suppose that 4.950 kg of ammonia, $NH_3(g)$, is available. What mass in grams of ammonium nitrate is produced, if the reaction is only 89.5% efficient?



What Is Required?

You need to calculate the mass in grams of ammonium nitrate, that forms in the reaction.

What Is Given?

You know the percentage yield of ammonium nitrate: 89.5%

You know the mass of ammonia: 4.950 kg

You know the balanced chemical equation:
$NH_3(g) + HNO_3(aq) \rightarrow NH_4NO_3(s)$

Plan Your Strategy	Act on Your Strategy
Convert the mass of ammonia from kilograms to grams.	$(4.950\ \cancel{kg}\ NH_3) \times \left(\frac{1000\ g}{1\ \cancel{kg}}\right) = 4.950 \times 10^3\ g\ NH_3$
Find the molar masses of ammonia and ammonium nitrate.	$M_{NH_3} = 1M_N + 3M_H = 1(14.01\ g/mol) + 3(1.01\ g/mol) = 17.04\ g/mol$ $M_{NH_4NO_3} = 2M_N + 4M_H + 3M_O$ $= 2(14.01\ g/mol) + 4(1.01\ g/mol) + 3(16.00\ g/mol) = 80.06\ g/mol$
Find the amount of ammonia in moles.	$n = \frac{m}{M} = \frac{4.950 \times 10^3\ g}{17.04\ g/mol}$ $= 290.493\ mol\ NH_3$
Calculate the theoretical yield of ammonium nitrate in moles.	$n_{NH_4NO_3} = 290.493\ \cancel{mol\ NH_3} \times \left(\frac{1\ mol\ NH_4NO_3}{1\ \cancel{mol\ NH_3}}\right)$ $= 290.493\ mol\ NH_4NO_3$
Calculate the theoretical yield of ammonium nitrate in grams.	$m = n \times M = 290.493\ \cancel{mol} \times 80.06\ g/\cancel{mol}$ $= 2.3257 \times 10^4\ g\ NH_4NO_3$
Use the percentage yield equation to solve for the unknown.	$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$ $\text{actual yield} = \frac{\text{percentage yield} \times \text{theoretical yield}}{100\%}$ $= \frac{(89.5\cancel{\%})(2.3257 \times 10^4\ g\ NH_4NO_3)}{100\cancel{\%}}$ $= 2.08 \times 10^4\ g\ NH_4NO_3$ Therefore, the actual yield of ammonium nitrate is $2.08 \times 10^4\ g\ NH_4NO_3$.

Check Your Solution

The actual yield should be about 90% of the theoretical yield, so the answer is reasonable.

Practice Problems

51. During an investigation, calcium carbide, $CaC_2(s)$, reacted with excess water to make calcium hydroxide, $Ca(OH)_2(aq)$, and acetylene, $C_2H_2(g)$.

$$CaC_2(s) + 2H_2O(\ell) \rightarrow Ca(OH)_2(aq) + C_2H_2(g)$$

The data table for this investigation is given below.

Data Table

Mass of Calcium Carbide That Reacted	2.38 g
Mass of Acetylene That Was Produced	0.77 g

What was the theoretical yield and actual yield of acetylene?

52. Suppose that 0.250 mol of potassium carbonate, $K_2CO_3(s)$, reacts with excess hydrochloric acid as follows:

$$K_2CO_3(s) + 2HCl(aq) \rightarrow H_2O(\ell) + CO_2(g) + 2KCl(aq)$$

a. Calculate the theoretical yield of potassium chloride.

b. Calculate the percentage yield of water if 0.189 mol of water is produced.

53. Phosphoric acid, $H_3PO_4(aq)$, is neutralized by potassium hydroxide, KOH(aq), according to the following reaction:

$$H_3PO_4(aq) + 3KOH(aq) \rightarrow K_3PO_4(aq) + 3H_2O(\ell)$$

If 49.0 g of potassium phosphate, $K_3PO_4(aq)$, is recovered after 49.0 g of phosphoric acid reacts with 49.0 g of potassium hydroxide, what is the percentage yield of the reaction?

54. The reaction of glucose, $C_6H_{12}O_6(s)$, with sulfuric acid, $H_2SO_4(\ell)$, produces carbon as follows:

$$C_6H_{12}O_6(s) + 2H_2SO_4(\ell) \rightarrow 6C(s) + 6H_2O(\ell) + 2H_2SO_4(aq)$$

a. If 20.8 g of glucose reacts with excess sulfuric acid, what is the theoretical yield, in grams, of carbon?

b. If the percentage yield is 72.0%, what mass of carbon is produced?

55. Calcium chloride, $CaCl_2(aq)$, is mixed with silver nitrate, $AgNO_3(aq)$, to form calcium nitrate, $Ca(NO_3)_2(aq)$, and silver chloride, AgCl(s).

$$CaCl_2(aq) + 2AgNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2AgCl(s)$$

If this reaction has an 81.5% yield, what mass of silver chloride is produced when 21.2 g of calcium chloride is added to excess silver nitrate?

56. The following reaction has a 68% yield.

$$AlCl_3(aq) + 4NaOH(aq) \rightarrow NaAlO_2(aq) + 3NaCl(aq) + 2H_2O(\ell)$$

Calculate the actual mass of sodium chloride that is recovered if 18.2 g of aluminum chloride, $AlCl_3(aq)$, reacts with 16.00 g of sodium hydroxide.

57. Ethyl butanoate, $C_6H_{13}O_2(\ell)$, is an organic ester that has the flavour and scent of pineapple. It is prepared as follows:

$$C_4H_9O_2(\ell) + C_2H_6O(\ell) \rightarrow C_6H_{13}O_2(\ell) + H_2O(\ell)$$

During an investigation, 0.573 mol of butanoic acid, $C_4H_9O_2(\ell)$, reacts with excess ethanol, $C_2H_6O(\ell)$. What mass of ethyl butanoate is produced if this reaction has a 92.0% yield?

58. An impure sample of barium hydroxide, $Ba(OH)_2(aq)$, has a mass 0.540 g. It is dissolved in water and then treated with excess sulfuric acid, $H_2SO_4(aq)$. This results in the formation of a precipitate of barium sulfate, $BaSO_4(s)$.

$$H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow BaSO_4(s) + 2H_2O(\ell)$$

The barium sulfate is filtered, and any remaining sulfuric acid is washed away. Then the barium sulfate is dried and its mass is measured to be 0.62 g. What mass of barium hydroxide was in the original (impure) sample?

59. Iron pyrite, $FeS_2(s)$, reacts with oxygen as shown in the reaction below:

$$4FeS_2(s) + 11O_2(g) \rightarrow 2Fe_2O_3(s) + 8SO_2(g)$$

a. In a laboratory, 5.000 kg of an impure mineral, which contains 45.3% iron pyrite, reacts with oxygen. Calculate the mass of iron(III) oxide, $Fe_2O_3(s)$, that forms. Assume that all the pyrite reacts.

b. Suppose that the reaction has a 78.0% yield, due to an incomplete reaction. How many grams of iron(III) oxide is produced?

60. Sodium oxide, $Na_2O(s)$, reacts with water to form the base sodium hydroxide.

$$Na_2O(s) + H_2O(\ell) \rightarrow 2NaOH(aq)$$

If this reaction has a 91% yield, what mass of sodium hydroxide is obtained when 0.483 mol of sodium oxide reacts with excess water?

Activity 7.4 Increasing Percentage Yield

The cost of manufactured items is primarily based on the cost of producing the items. Many manufacturers aim to produce items as inexpensively as possible, while maintaining the quality of the items. They try to ensure that they use enough reactants to produce a high yield. At the same time, they must try to ensure that they do not waste reactants or introduce chemical waste into the environment. Wasting reactants unnecessarily increases the cost of producing the items. Introducing chemical waste into the environment can create environmental damage that is expensive to clean up and can sometimes not be undone.

Suppose that you work for a company that wants to test iron(III) oxide, $Fe_2O_3(s)$, as a pigment for cosmetics. You want to produce iron(III) oxide from steel wool.

Procedure

Note any visible differences you can see in the two different grades of steel wool in the photographs below, and then answer the questions.

Questions

1. Steel wool contains iron and small amounts of carbon. If the steel wool is held over a Bunsen burner, the iron in the steel wool combines with oxygen in the air and forms iron(III) oxide. Write a balanced chemical equation for this reaction.
2. How would you determine the actual and theoretical percentage yields of iron(III) oxide for each grade of steel wool?
3. While conducting the reaction described in question 1, you discover that oxygen is a limiting factor. What could you do to overcome this limiting factor and increase your percentage yield?
4. Suppose that you can buy steel wool from two different companies. The steel wool from one company contains less iron and more impurities. The steel wool from the other company is almost pure iron but is more expensive. Describe at least two factors that you must consider before choosing which steel wool you will use to produce your iron(III) oxide. Explain your reasoning.

Suggested Investigation

Inquiry Investigation 7-C, Finding the Percentage Yield of a Double Displacement Reaction

The Importance of Actual Percentage Yield

You have learned how to predict the amount of product that forms from a certain amount of reactant. It is important to maximize the percentage yield of any chemical process, but there are several reasons why the maximum yield may not be produced. The amount that is predicted by stoichiometric calculations may be reduced because of impure reactants or because of how the product is collected.

It is often useful to know the percentage yield of a chemical reaction. If you know that the percentage yield of a reaction is about 14 percent, you can calculate the amounts of reactants that are needed to produce a desired mass of product. It is important to know the percentage yield of a large chemical process because of the cost of production. A small difference in the percentage yield of a product could mean a loss of thousands of dollars for a chemical company.

Section 7.3 REVIEW

Section Summary

- Stoichiometric calculations are used to determine the theoretical yield, which is the maximum theoretical amount of product that can be produced from a given amount of reactants in a chemical reaction.
- The actual yield is determined through experimentation and is the actual amount of product that is produced in a chemical reaction.
- The percentage yield is the ratio of the actual yield to the theoretical yield, expressed as a percent.

Review Questions

1. **K/U** Does it matter which units you use, grams or moles, when calculating the percentage yield of a reaction? Explain.
2. **K/U** Briefly compare and contrast the terms "theoretical yield" and "actual yield" using an analogy of your own choice.
3. **K/U** Explain why the reaction rate can reduce the actual yield of product.
4. **T/I** Silver nitrate, $AgNO_3(aq)$, reacts with potassium bromide, $KBr(aq)$, to produce silver bromide, $AgBr(s)$.
   $$AgNO_3(aq) + KBr(aq) \rightarrow AgBr(s) + KNO_3(aq)$$
   If 14.64 g of silver bromide is obtained when 14.00 g of silver nitrate reacts with excess potassium bromide, what is the percentage yield?
5. **T/I** Sodium phosphate, $Na_3PO_4(aq)$, reacts with barium nitrate, $Ba(NO_3)_2(aq)$, as represented by:
   $$2Na_3PO_4(aq) + 3Ba(NO_3)_2(aq) \rightarrow Ba_3(PO_4)_2(s) + 6NaNO_3(aq)$$
   If 5.00 g of sodium phosphate reacts with 10.90 g of barium nitrate and 7.69 g of solid precipitate is recovered, what is the percentage yield?
6. **K/U** Zinc reacts with sulfuric acid, $H_2SO_4(aq)$, to yield zinc sulfate, $ZnSO_4(aq)$, and hydrogen gas. What assumption is usually made about the purity of the reactants, such as the zinc?
7. **C** Use a flowchart to describe the steps you would take to find the percentage yield of a reaction, given the mass of the two reactants and the mass of the product.
8. **C** You find a website that describes an investigation to produce carbon dioxide gas using the following equation:
   $$CH_3COOH(aq) + NaHCO_3(s) \rightarrow NaCH_3COO(aq) + CO_2(g) + H_2O(\ell)$$
   The procedure states that household white vinegar, which is dilute acetic acid, $CH_3COOH(aq)$, could be used instead of pure acetic acid. Write a posting for the site explaining why the product yield will be low when vinegar is used instead of pure acetic acid.
9. **A** Explain why percentage yield is important to companies that produce chemical compounds for profit.
10. **A** Explain how each laboratory technique given below could be changed to increase the yield.
    a. stirring the reaction material with a wooden splint
    b. transferring a reactant from one beaker to another
    c. drying a mixture in an evaporating dish over a laboratory gas burner
    d. adding aqueous reactants from a graduated cylinder to the reaction beaker
11. **T/I** Sodium iodide, $NaI(s)$, and chlorine gas undergo a single displacement reaction, as represented by the following equation:
    $$2NaI(s) + Cl_2(g) \rightarrow I_2(s) + 2NaCl(s)$$
    a. If 4.0 g of sodium iodide and 4.0 g of chlorine gas react, what are the theoretical yields of both products?
    b. If the percentage yield of sodium chloride is 67%, what is the actual yield of sodium chloride?
    c. Would it be appropriate to assume that the yield of the other product, iodine, $I_2(s)$, is also 67%? Explain.
12. **T/I** To dehydrate bluestone, which is a form of copper(II) sulfate pentahydrate, $CuSO_4 \cdot 5H_2O(s)$, it was heated over a Bunsen burner. The following equation represents the dehydration reaction:
    $$CuSO_4 \cdot 5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(\ell)$$
    Three masses were measured, before and after the reaction, as shown in the data table below. Use these masses to calculate the theoretical yield, actual yield, and percentage yield of both copper sulfate, $CuSO_4(s)$, and water.

Data Table

Mass of Empty Crucible and Lid	15.146 g
Mass of Crucible and Copper(II) Sulfate Pentahydrate	19.273 g
Mass of Crucible and Copper Sulfate	18.059 g