7.5 Percentage Yield

Winemaking is one of humankind's oldest technologies based on chemical reactions. Evidence of the earliest winemakers dates back to about 6500 years ago. A key chemical process in the production of wine is fermentation (**Figure 1**). Fermentation is a complicated process during which yeast cells metabolize sugars such as glucose, $C_6H_{12}O_6$, in grape juice. The products of this reaction are an alcohol called ethanol, C_2H_5OH, and carbon dioxide:

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

The toxicity of the ethanol limits the extent to which fermentation can occur. As the concentration of ethanol in the mixture increases, yeast cells begin to die. This stops fermentation before all of the glucose is converted into ethanol. As a result, the volume of ethanol in a bottle of wine rarely exceeds 14 % of the overall volume.

Figure 1 Fermentation locks filled with water are fitted to the top of each bottle, allowing carbon dioxide to escape while preventing oxygen from entering. If the bottles were sealed, the pressure from the accumulating carbon dioxide could break them.

Reaction Yield

Imagine a chemical reaction with a limiting reagent. Ideally, all of the limiting reagent is converted into the desired product. This gives the maximum possible yield of product and is defined as the theoretical yield for the reaction. The **theoretical yield** is a prediction of how much product should form based on the stoichiometry of the reaction. The theoretical yield is achieved if 100 % of the limiting reagent is converted into products. Chemists can predict the theoretical yield before observing the reaction. You predicted theoretical yields in Section 7.4, when you calculated the amount or the mass of product that could be produced.

In practice, however, theoretical yields are rarely achieved. Instead, the **actual yield**—the amount of the product collected during an experiment or industrial process—is usually less. This means that some of the limiting reagent did not become part of the collected product.

There are several factors that can account for why the actual yield is usually less than the theoretical yield. We will look at four of these factors in some detail.

theoretical yield the amount or mass of product predicted based on the stoichiometry of the chemical equation

actual yield the amount or mass of product actually collected during an experiment or industrial process

The Nature of the Reaction

The complete conversion of reactants to products is not always possible. This is sometimes because a competing reverse reaction takes place. For example, ammonia, $NH_3(g)$, is an important industrial chemical made from hydrogen and nitrogen:

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

As the reaction proceeds, ammonia begins to decompose (**Figure 2**):

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

Chemists have determined the temperature, pressure, and catalyst that give the best yield of ammonia. However, the actual yield is never equal to the theoretical yield.

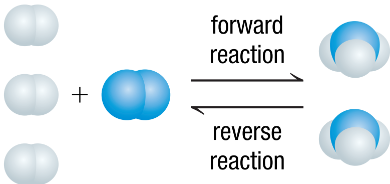

Figure 2 Some ammonia spontaneously decomposes as it forms. This reverse reaction limits the yield of ammonia.

The Experimental Procedure

Regardless of how careful chemists are, they inevitably lose small quantities of material in most investigations. If some of the mass of a reactant is lost after it is measured, or some of the mass of the product is lost before it is measured, the actual yield will be less than the theoretical yield. Materials are lost through spillage during the transfer of solutions, by splattering during heating, and losses during the isolation and purification of the product. Chemists can reduce these losses by improving their skills, using better equipment, and reducing the number of steps in the procedure. Having fewer steps reduces the chances of accidental loss of product.

Impurities

The chemicals used in investigations are rarely 100 % pure. In fact, they come in a wide range of grades (or purities). The mass of the reactant in any sample of starting chemical is therefore less than the measured mass. Failure to take into account the impurities when determining the theoretical yield results in a predicted amount of product that is impossible to achieve. As chemicals age, they may become more impure—particularly if they are not stored properly. Sodium hydroxide, for example, must be stored in an airtight container because it readily absorbs water. Magnesium forms an oxide layer if it is exposed to the air. This increases the mass of a strip of "magnesium" about to be used in an investigation.

Competing Side Reactions

Sometimes, competing reactions prevent some of the reactants from being converted into products. For example, during the synthesis of cisplatin, the chemotherapy drug introduced in Section 7.3, a similar compound called transplatin is also formed (**Figure 3**). Note that cisplatin and transplatin differ only in the position of the groups of atoms attached to the central platinum atom. At first, this difference does not appear to be significant. However, researchers believe that this difference accounts for why transplatin has little to no effect as a cancer treatment. Chemists have found ways to minimize the formation of transplatin and maximize the yield of cisplatin. How they do this, though, is beyond the scope of this course.

Cisplatin Transplatin

Figure 3 Cisplatin is an important drug in the treatment of cancer. The yield of cisplatin is reduced by a competing reaction that produces transplatin.

Percentage Yield

Comparing the actual yield in a reaction to the theoretical yield gives an indication of how efficient or successful the reaction is at converting reactants into products. A numerical value of this efficiency is determined by calculating the percentage yield for the reaction (**Figure 4**). **Percentage yield** is defined as

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

Percentage yield can be calculated either by comparing the actual mass of product with the theoretical mass, or by comparing the actual amount of product with the theoretical amount.

Chemists often report the percentage yield when they synthesize a new compound or develop a new way of making an existing compound. This information gives other chemists an idea of how much product to expect, if they repeat the procedure.

percentage yield the ratio, expressed as a percentage, of the actual yield to the theoretical yield

Figure 4 Only 16 of the initial 20 kernels of popping corn popped. The percentage yield of popcorn is $\frac{16}{20} \times 100\ \% = 80\ \%$.

Tutorial 1 Determining Percentage Yield

In a typical percentage yield problem, you are given the masses of reactants and the mass of a desired product.

Here are some helpful tips to keep in mind when solving the problem:

- Use stoichiometry (based on the balanced chemical equation) to determine the theoretical yield of the desired product.
- If you are told which of the reactants the limiting reagent is, use its mass to calculate theoretical yield (Section 7.2). If the masses of both reactants are given, you must first determine the limiting reagent (Sections 7.3 and 7.4).
- Make sure that the actual yield and theoretical yield are expressed in the same units: either grams or moles. Note that percentage yield has no units because the units of actual and theoretical yield cancel each other in the calculation.

Sample Problem 1 follows essentially the same steps that you used in Section 7.2. The only difference is the calculation of percentage yield in the last step.

Sample Problem 1: Determining Percentage Yield from Mass Data

Methanol, CH_3OH, can be made in a synthesis reaction using carbon dioxide and hydrogen (**Figure 5**):

$$CO_2(g) + 3\ H_2(g) \rightarrow CH_3OH(l) + H_2O(g)$$

During an investigation, 20.0 g of hydrogen was reacted with excess carbon dioxide to produce 102.0 g of methanol. What is the percentage yield of this reaction?

WEB LINK

For online tutorials on determining percentage yield,

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Figure 5 Some high-performance race cars use methanol as fuel.

Given: $m_{H_2} = 20.0$ g; $m_{CH_3OH} = 102.0$ g; carbon dioxide is in excess

Required: percentage yield of methanol

Solution:

Step 1. Write a balanced equation listing given value(s), required value(s), and corresponding molar masses.

$$\begin{array}{llll} CO_2(g) + & 3\,H_2(g) \rightarrow & CH_3OH(l) + & H_2O(g) \\ & 20.0\text{ g} & 102.0\text{ g} & \\ & 2.02\text{ g/mol} & 32.05\text{ g/mol} & \end{array}$$

Step 2. Convert mass of given substance to amount of given substance.

$$n_{H_2} = 20.0\,\cancel{\text{g}} \times \frac{1\text{ mol}}{2.02\,\cancel{\text{g}}}$$

$$n_{H_2} = 9.9010\text{ mol [extra digits carried]}$$

Step 3. Convert amount of given substance to amount of required substance.

$$n_{CH_3OH} = 9.9010\,\cancel{\text{mol}_{H_2}} \times \frac{1\text{ mol}_{CH_3OH}}{3\,\cancel{\text{mol}_{H_2}}}$$

$$n_{CH_3OH} = 3.3003\text{ mol}$$

Step 4. Convert amount of required substance to mass of required substance.

$$m_{CH_3OH} = (3.3003\,\cancel{\text{mol}})\left(\frac{32.05\text{ g}}{1\,\cancel{\text{mol}}}\right)$$

$$m_{CH_3OH} = 105.78\text{ g [extra digits carried]}$$

Step 5. Calculate the percentage yield.

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

$$= \frac{102.0\,\cancel{\text{g}_{CH_3OH}}}{105.78\,\cancel{\text{g}_{CH_3OH}}} \times 100\ \%$$

$$\text{percentage yield} = 96.4\ \%$$

Statement: The percentage yield in the synthesis of methanol reaction is 96.4 %.

Practice

SKILLS HANDBOOK A6

1. Impure nickel can be purified by first reacting it with carbon monoxide:

   $Ni(s) + 4\,CO(g) \rightarrow Ni(CO)_4(g)$ T/I

   (a) Calculate the theoretical yield of nickel carbonyl, $Ni(CO)_4$, in this reaction if 23.5 g of nickel reacts with excess carbon monoxide. [ans: 68.4 g]

   (b) Calculate the percentage yield if 61.0 g of nickel carbonyl is collected. [ans: 89.2 %]

2. An industrial process produces hydrogen by reacting methane with water vapour:

   $CH_4(g) + H_2O(g) \rightarrow CO(g) + 3\,H_2(g)$

   1.61 g of methane is combined with 2.00 g of water, producing 0.50 g of hydrogen. Determine the theoretical yield and percentage yield of hydrogen. T/I [ans: 0.608 g; 82 %]

7.5 Summary

- The theoretical yield is the quantity of product as predicted by the stoichiometry of the chemical equation.
- The actual yield of a reaction is the actual quantity of product collected.
- The actual yield is usually less than the theoretical yield due to material losses, the nature of the reaction, impurities, and competing side reactions.
- The percentage yield of the reaction is the ratio, expressed as a percentage, of the actual yield and theoretical yield. The value of percentage yield has no units.

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

Investigation 7.5.1

What Stopped the Silver? (p. 342)
In this experiment you will investigate the effect that the mass of silver nitrate has on the percentage yield of pure silver. You will have to decide which reactant is the limiting reagent.

7.5 Questions

1. How do actual yield and theoretical yield differ? K/U
2. Why are actual yield and theoretical yield rarely equal? K/U
3. Why does reducing the number of steps in an experimental procedure often result in an increase in percentage yield? K/U
4. A student carefully neutralized a sample of hydrochloric acid with sodium hydrogen carbonate in an open container:

   $HCl(aq) + NaHCO_3(s) \rightarrow H_2O(l) + CO_2(g) + NaCl(aq)$

   When the reaction was complete the student heated the remaining solution to remove the water. What effect, if any, would each of the following have on the percentage yield of sodium chloride? Justify your answer in each case. T/I
   (a) not drying the mixture long enough
   (b) adding insufficient sodium hydrogen carbonate to completely react all of the acid
   (c) using an impure form of sodium hydrogen carbonate
   (d) splattering of the reaction mixture out of the container during the reaction
5. Methane reacts with chlorine to form chloromethane:

   $CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g)$

   However, when the reaction mixture is analyzed, some dichloromethane, $CH_2Cl_2(g)$, is detected as well. What effect does this have on the percentage yield of chloromethane? Why? K/U T/I
6. A chemist adds 3.00 g of zinc to a solution containing an excess of silver nitrate. Only 7.2 g of silver metal is collected by the end of the investigation. T/I C
   (a) Write a balanced chemical equation for the reaction.
   (b) Determine the theoretical yield of silver metal.
   (c) Determine the percentage yield of this reaction.
7. To manufacture large quantities of chlorine, sodium hydroxide, and hydrogen, industrial technicians pass electricity through a concentrated sodium chloride solution:

$$2\ NaCl(aq) + 2\ H_2O(l) \xrightarrow{\text{electricity}} Cl_2(g) + 2\ NaOH(aq) + H_2(g)$$ T/I

   (a) If the percentage yield of chlorine in this reaction is 95 % and the theoretical yield is 142 g of chlorine, what is the actual yield of chlorine? (Assume that there is excess water present.)
   (b) What mass of sodium chloride is required to produce this yield of chlorine?
8. Coal gasification is a process that converts coal (mostly carbon) into methane, $CH_4(g)$:

   $2\ C(s) + 2\ H_2O(g) \rightarrow CH_4(g) + CO_2(g)$

   If 4.00 mol of coal is combined with excess water vapour, producing 28.0 g of methane, what is the percentage yield of the reaction? T/I
9. Ammonium sulfate, a fertilizer, is made using the following chemical reaction:

   $2\ NH_3(g) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq)$ T/I

   (a) What is the theoretical yield of ammonium sulfate (in grams) from 3.4 g of ammonia and excess sulfuric acid?
   (b) Determine the percentage yield if 10.4 g of ammonium sulfate is collected.
10. Iron is extracted from the mineral magnetite, Fe_3O_4:

    $Fe_3O_4(s) + 2\ C(s) \rightarrow 3\ Fe(s) + 2\ CO_2(g)$

    A 35.0 g sample of magnetite produces 15.0 g of iron. Determine the percentage yield of this reaction. T/I
11. At 300 °C and 25 times atmospheric pressure, the percentage yield for the production of ammonia from its elements is 24.5 %. If 1.00 kg of nitrogen reacts with excess hydrogen,
    (a) write the balanced chemical equation for this reaction.
    (b) what is the actual yield of ammonia?
    (c) what amount of nitrogen remains unreacted?
    (d) what mass of nitrogen remains unreacted? T/I C
12. Ethanoic acid can be produced industrially from the reaction of methanol, $CH_3OH(g)$, with carbon monoxide, $CO(g)$:

    $CH_3OH(l) + CO(g) \rightarrow HC_2H_3O_2(l)$

    The percentage yield of this reaction is known to be 85 %. What mass of carbon monoxide is required to prepare 2.5 kg of ethanoic acid if methanol is present in excess? T/I
13. The percentage yield of the following reaction is consistently 75.2 %.

    $CH_4(g) + 4\ S(g) \rightarrow CS_2(g) + 2\ H_2S(g)$

    What mass of sulfur is required to produce 50.0 g of hydrogen sulfide if methane, CH_4, is present in excess? T/I
14. Phosphoric acid, $H_3PO_4(aq)$, is used in rust removal products and soft drinks. Phosphoric acid can be made in a two-step process:

    $P_4(s) + 5\ O_2(g) \rightarrow P_4O_{10}(s)$

    $P_4O_{10}(s) + 6\ H_2O(l) \rightarrow 4\ H_3PO_4(aq)$

    In the first step, 25.0 g phosphorus burns in excess oxygen to produce tetraphosphorus decoxide. The percentage yield of this reaction is 90.0 %. T/I
    (a) What mass of tetraphosphorus decoxide is collected from the first step?
    (b) The percentage yield of the second reaction, in excess water, is 95.2 %. What mass of phosphoric acid is produced?
15. Research "atom economy" and prepare a one-paragraph summary of how it differs from percentage yield. K/U T/I A

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