8.7: Theoretical Yield and Percent Yield

Learning Objectives

- Calculate percentage or actual yields from known amounts of reactants.

The world of pharmaceutical production is an expensive one. Many drugs have several steps in their synthesis and use costly chemicals. A great deal of research takes place to develop better ways to make drugs faster and more efficiently. Studying how much of a compound is produced in any given reaction is an important part of cost control.

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 predicted. Besides spills and other experimental errors, there are usually losses due to an incomplete reaction, undesirable side reactions, etc. 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. As we have previously learned, this is called the **theoretical yield**, the maximum amount of product that can 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. Once again,

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

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 unnecessary expense.

The percent yield of a desired product may never exceed 100% for the reasons indicated earlier. Should a percent yield ever be greater than 100%, this would be an indication of some sort of experimental error or that the reaction contains impurities that cause its mass to be greater than if the product was pure. When chemists synthesize a desired chemical, they are always careful to purify the products of the reaction. Example 8.7.1 illustrates the steps for determining percent yield.

Example 8.7.1: Theoretical Yield and Percent Yield

Potassium chlorate decomposes upon heating, according to the reaction below:

$$2\ KClO_3\ (s) \xrightarrow{\triangle} 2\ KCl\ (s) + 3\ O_2\ (g)$$

If 40.0 g $KClO_3$ is heated until it completely decomposes, resulting in the collection of 14.3 g of oxygen gas,

A. What is the theoretical yield of oxygen gas?
B. What is the percent yield of oxygen gas?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 40.0 g $KClO_3$ reacted; 14.3 g O_2 produced Find: theoretical yield O_2; percent yield O_2
List other known quantities.	1 mol $KClO_3$ = 122.55 g $KClO_3$ 1 mol O_2 = 32.00 g O_2 2 mol $KClO_3$: 3 mol O_2
Prepare concept maps using the proper conversion factor(s).	$\boxed{\text{g } KClO_3} \xrightarrow[\text{122.55 g } KClO_3]{\text{1 mol } KClO_3} \boxed{\text{mol } KClO_3} \xrightarrow[\text{2 mol } KClO_3]{\text{3 mol } O_2} \boxed{\text{mol } O_2} \xrightarrow[\text{1}]{\text{32.0}}$
Calculate the theoretical yield.	$40.0\ \text{g}\ \cancel{KClO_3} \times \frac{1\ \cancel{\text{mol } KClO_3}}{122.55\ \cancel{\text{g } KClO_3}} \times \frac{3\ \cancel{\text{mol } O_2}}{2\ \cancel{\text{mol } KClO_3}} \times \frac{32.00\ \text{g } O_2}{1\ \cancel{\text{mol } O_2}} = \boxed{15.7\ \text{g } O_2}$
Calculate the percent yield.	$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{14.3\ \cancel{\text{g } O_2}}{15.7\ \cancel{\text{g } O_2}} \times 100\% = \boxed{91.1\%\ \text{yield}}$
Think about your result.	The percent yield is less than 100%, but still quite high (over 90%). This seems like a reasonable yield for a chemical reaction.

Example 8.7.2: Limiting Reactants, Theoretical Yield, and Percent Yield

Vanadium metal reacts with oxygen gas to yield solid vanadium(V) oxide:

$$4\ V\ (s) + 5\ O_2\ (g) \rightarrow 2\ V_2O_5\ (s)$$

A. What mass of V_2O_5 can be made when 41.3 g V reacts with 35.0 g O_2?
B. If 62.3 g V_2O_5 were actually produced, what is the percent yield?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 41.3 g V reacted; 35.0 g O_2 reacted; 62.3 g V_2O_5 produced Find: theoretical yield V_2O_5; percent yield V_2O_5
List other known quantities.	1 mol V = 50.94 g V 1 mol O_2 = 32.00 g O_2 1 mol V_2O_5 = 181.88 g V_2O_5 4 mol V: 2 mol V_2O_5 5 mol O_2: 2 mol V_2O_5
Prepare concept maps using the proper conversion factor(s).	$\boxed{\text{g V}} \xrightarrow[\text{50.94 g V}]{\text{1 mol V}} \boxed{\text{mol V}} \xrightarrow[\text{4 mol V}]{\text{2 mol } V_2O_5} \boxed{\text{mol } V_2O_5} \xrightarrow[\text{1 mol } V_2O_5]{\text{181.88 g } V_2O_5} \boxed{\text{g V}}$ $\boxed{\text{g } O_2} \xrightarrow[\text{32.00 g } O_2]{\text{1 mol } O_2} \boxed{\text{mol } O_2} \xrightarrow[\text{5 mol } O_2]{\text{2 mol } V_2O_5} \boxed{\text{mol } V_2O_5} \xrightarrow[\text{1 mol } V_2O_5]{\text{181.88 g } V_2O_5} \boxed{\text{g}}$
Calculate the theoretical yield. Select the smallest answer.	$41.3 \text{ g } \cancel{V} \times \frac{1 \cancel{\text{ mol V}}}{50.94 \text{ g } \cancel{V}} \times \frac{2 \text{ mol } \cancel{V_2O_5}}{4 \cancel{\text{ mol V}}} \times \frac{181.88 \text{ g } V_2O_5}{1 \text{ mol } \cancel{V_2O_5}} = \boxed{73.7 \text{ g } V_2O_5}$ $35.0 \text{ g } \cancel{O_2} \times \frac{1 \text{ mol } \cancel{O_2}}{32.00 \text{ g } \cancel{O_2}} \times \frac{2 \text{ mol } \cancel{V_2O_5}}{5 \text{ mol } \cancel{O_2}} \times \frac{181.88 \text{ g } V_2O_5}{1 \text{ mol } \cancel{V_2O_5}} = \cancel{79.6 \text{ g } V_2O_5}$
Calculate the percent yield.	$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{62.3 \text{ g } \cancel{V_2O_5}}{73.7 \text{ g } \cancel{V_2O_5}} \times 100\% = \boxed{84.5\% \text{ yield}}$
Think about your result.	The percent yield is less than 100%, but still quite high (almost 85%). This seems like a reasonable yield for a chemical reaction.

Exercise 8.7.1

What is the percent yield of a reaction that produces 12.5 g of the Freon CF_2Cl_2 from 32.9 g of CCl_4 and an excess of HF?

$$CCl_4 + 2\,HF \rightarrow CF_2Cl_2 + 2\,HCl$$

Answer

48.3% yield

Summary

- Theoretical yield is calculated based on the stoichiometry of the chemical equation.
- The actual yield is experimentally determined.
- The percent yield is determined by calculating the ratio of actual yield to theoretical yield.