Stoichiometry

Chapter 11

Objectives

- **Describe** the types of relationships indicated by a balanced chemical equation
- **State** the mole ratios from a balanced chemical equation
- **MAIN IDEA:** The amount of each reactant present at the start of a chemical reaction determines how much product can be formed

Stoichiometry

- **Stoichiometry** – study of the amount of substances consumed and produced in chemical reactions
  - Can calculate the quantities of substances in a chemical equation from other substances as long as the chemical equation is **balanced**
  - Stoichiometry will combine all of the things we have already learned
  - Chemical reactions stop when one of the reactants is used up

- **To solve stoich problems**:
  - Will need a **balanced** chemical equation
  - Determine known(s) and unknown(s)
  - Determine mole ratio of known and unknown(s)

Remember the 2 clear liquids that formed a yellow precipitate?

- 4.0 grams of Potassium iodide is combined with an excess amount lead II nitrate will produce how many grams of precipitate?

1st write Balanced equation

2nd include knowns and unknowns

$2KI_{(aq)} + Pb(NO_3)_{2\ (aq)} \rightarrow 2KNO_{3\ (aq)} + PbI_{2\ (s)}$

4.0 grams **? grams**

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- The **coefficients** in the balanced eqn give the relative number of moles of reactants and products
- A **molar ratio** is a ratio between the numbers of moles of any two substances in a balanced equation. **Gets you from what you know to what you don't know!**
- In Stoichiometric Calculations – **Must go through MOLES!!!** Will use MOLAR RATIO
- How do we start problem?? START WITH WHAT YOU KNOW!!!!!!

Molar Ratio

$$\left(\frac{4.0\ g\ KI}{1}\right)\left(\frac{1\ mole\ KI}{166.0\ g\ KI}\right)\left(\frac{1\ mole\ PbI_2}{2\ mole\ KI}\right)\left(\frac{461.0\ g\ PbI_2}{1\ mole\ PbI_2}\right) = \boxed{5.6\ g\ PbI_2}$$

Objectives – Limiting Reactants

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

Why do reactions stop?

- Reactions proceed until one of the reactants is **used up** and one is left in **excess**
- The **Limiting Reactant** limits the extent of the reaction and determines the amount of product formed
- The **Excess Reactant** are all the leftover unused reactants
- Determining the limiting reactant is important **because the amounts of the products formed depends on this reactant**

Cookie Recipe

Supply in House (given)	Recipe (formula)	Product
12 eggs	2 eggs	1 dozen cookies
20 cups of flour	2 cups flour	
3 cups of chips	1 cup chips	
5 cups sugar	1 cup sugar	

- Which will you run out of 1St ?
  - Chips: This is the **Limiting Reactant**
  - The rest are Excess Reactants
- How much product can we make?
  - Chips is limiting so you can make 3 dozen cookies

Limiting/Excess Reactions

- It is a **limiting**/**excess** problem when the amounts of more than one reactant are given

Which one is Excess?

$$2Mg_{(s)} + O_{2(g)} \rightarrow 2\,MgO_{(s)}$$

How much Excess Remains?

Have 5.6 g 15.6 g ? g

How many grams of MgO are produced?

Step 1: Ignore amount Given for one of the Reactants

Determine how much you **NEED** of Ignored Reactant

$$\left(\frac{5.6\text{ g Mg}}{1}\right)\left(\frac{1\text{ mole Mg}}{24.3\text{ g Mg}}\right)\left(\frac{1\text{ mole }O_2}{2\text{ mole Mg}}\right)\left(\frac{32.0\text{ g }O_2}{1\text{ mole }O_2}\right) = \boxed{3.7\text{ g }O_2}$$

Need

$$2Mg_{(s)} + O_{2(g)} \rightarrow 2\,MgO_{(s)}$$

Have 5.6 g 15.6 g ? g

Need 3.7 grams O_2

Step 2: Compare what you **HAVE** and what you **NEED**

Determine Excess and Limiting

15.6 g O2 Have

3.7 g O2 Need

11.9 g O2 Excess

And then Mg has to be the limiting reactant

$$\underset{5.6\ g}{2Mg_{(s)}} + \underset{15.6\ g}{O_{2\,(g)}} \rightarrow \underset{?\ g}{2\ MgO_{(s)}}$$

Limiting Excess

Step 3: Determine Products – START with LIMITING!!!!!

$$\left(\frac{5.6\text{ g Mg}}{1}\right)\left(\frac{1\text{ mole Mg}}{24.3\text{ g Mg}}\right)\left(\frac{2\text{ mole MgO}}{2\text{ mole Mg}}\right)\left(\frac{40.3\text{ g MgO}}{1\text{ mole MgO}}\right) = \boxed{9.3\ \text{g MgO}}$$

How much Product?

- Laboratory reactions do not always produce the calculated amounts of products
  - Reactants stick to containers
  - Spills
  - Competing reactions form other by products
- The **Theoretical yield** is the maximum amount of product that can be produced from a given amount of reactant
  - This is what we figure out with **math**!
- The **actual yield** is the amount of product actually produced when the chemical reaction is carried out in an experiment
  - Accounts for lost chemicals: The **amount measured** in the lab!

How much product?

- The **percent yield** of a product is the ratio of the actual yield expressed as a percent
  - Percent yield is important in the cost effectiveness of many industrial manufacturing processes

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

Key Concepts

- Balanced chemical equations can be interpreted in terms of moles, mass, and representative particles (atoms, molecules, formula units).
- The **law of conservation of mass** applies to all chemical reactions.
- **Mole ratios** are derived from the coefficients of a balanced chemical equation. Each mole ratio relates the number of moles of one reactant or product to the number of moles of another reactant or product in the chemical reaction
- Chemists use stoichiometric calculations to predict the amounts of reactants used and products formed in specific reactions.
- The first step in solving stoichiometric problems is writing the balanced chemical equation.
- Mole ratios derived from the balanced chemical equation are used in stoichiometric calculations.

Key Concepts

- Stoichiometric problems make use of mole ratios to convert between mass and moles
- The limiting reactant is the reactant that is completely consumed during a chemical reaction. Reactants that remain after the reaction stops are called excess reactants.
- To determine the limiting reactant, the actual mole ratio of the available reactants must be compared with the ratio of the reactants obtained from the coefficients in the balanced chemical equation.
- Stoichiometric calculations must be based on the limiting reactant.

Key Concepts

- The **theoretical yield** of a chemical reaction is the maximum amount of product that can be produced from a given amount of reactant. Theoretical yield is calculated from the balanced chemical equation.
- The **actual yield** is the amount of product produced. Actual yield must be obtained through experimentation.
- **Percent yield** is the ratio of actual yield to theoretical yield expressed as a percent. High percent yield is important in reducing the cost of every product produced through chemical processes.