123 Limiting Reagent And Percent Yield Answers

Limiting Reagents and Percentage Yield Worksheet

1. Consider the reaction

$I_2O_5(g) + 5\ CO(g)$ -------> $5\ CO_2(g) + I_2(g)$

a) 80.0 grams of iodine(V) oxide, I_2O_5, reacts with 28.0 grams of carbon monoxide, CO. Determine the mass of iodine I_2, which could be produced?

80 g I_2O_5	1 mol I_2O_5	1 mol I_2		XS
1	333.8 g I_2O_5	1 mol I_2O_5		

28 g CO	1 mol CO	1 mol I_2	253.8 g I_2	= 50.8 g I_2
1	28 g CO	5 mol CO	1 mol I_2	

b) If, in the above situation, only 0.160 moles, of iodine, I_2 was produced.

i) what mass of iodine was produced?

0.160 mol I_2	253.8 g I_2	= 40.6 g I_2
	1 mol I_2	

ii) what percentage yield of iodine was produced.

$$\frac{40.6\ g\ I_2}{50.8\ g\ I_2} \times 100 = 79.9\ \%$$

2. Zinc and sulphur react to form zinc sulphide according to the equation.

$Zn + S \rightarrow ZnS$

If 25.0 g of zinc and 30.0 g of sulphur are mixed,

a) Which chemical is the limiting reactant?

25.0 g Zn	1 mol Zn	1 mol ZnS	= 382 mol ZnS	LR
1	65.41 g Zn	1 mol Zn		

30.0 g S	1 mol S	1 mol ZnS	=934mol ZnS	XS
1	32.1g S	1 mol S		

b) How many grams of ZnS will be formed?

382 mol ZnS	97.51 g ZnS	= 37.2 g ZnS
	1 mol ZnS	

c) How many grams of the excess reactant will remain after the reaction is over?

.934mol ZnS -.382 mol ZnS= .552mol ZnS

552 mol ZnS	1 mol S	32.1g S	= 17.7g S
	1 mol ZnS	1 mol S	

3.

123 limiting reagent and percent yield answers are fundamental concepts in the field of chemistry, particularly when performing stoichiometric calculations in chemical reactions. Understanding these concepts is crucial for predicting the outcomes of reactions, ensuring efficient use of reactants, and evaluating the success of a reaction based on the amount of product produced. This article will delve into the definitions, significance, and methods to calculate the limiting reagent and percent yield in chemical reactions, providing examples for clarity.

Understanding Limiting Reagent

Definition

The limiting reagent, also known as the limiting reactant, is the substance that is entirely consumed first in a chemical reaction, thus limiting the amount of product that can be formed. In any given reaction, reactants are combined in specific ratios according to their stoichiometric coefficients. When one reactant is present in a lesser amount than is required for complete reaction with the other reactants, it becomes the limiting reagent.

Significance of Limiting Reagent

Identifying the limiting reagent is essential for several reasons:

1. Predicting Product Formation: Knowing which reactant will be consumed first allows chemists to predict how much product can be formed from the available reactants.
2. Cost Efficiency: By optimizing the use of reactants, chemists can minimize waste and reduce costs in industrial applications.
3. Reaction Yield: It helps in calculating the theoretical yield of products, which is necessary for determining percent yield.

Calculating the Limiting Reagent

To determine the limiting reagent in a reaction, follow these steps:

1. Write the Balanced Equation: Ensure the chemical equation is balanced. For example:
$$
aA + bB \rightarrow cC + dD
$$

2. Convert Units to Moles: If the amounts of reactants are given in grams or liters, convert them to moles using molar mass or molarity.

3. Calculate the Mole Ratios: Using the coefficients from the balanced equation, calculate how many moles of each reactant are required for the reaction.

4. Compare Available Moles to Required Moles: Determine which reactant will run out first based on the calculated ratios.

Example Calculation

Consider the reaction of hydrogen gas (H_2) with nitrogen gas (N_2) to form ammonia (NH_3):

$$
3H_2 + N_2 \rightarrow 2NH_3
$$

Suppose we have:
- 5 moles of H_2
- 2 moles of N_2

Step 1: From the balanced equation, 3 moles of H_2 are needed for every 1 mole of N_2.

Step 2: Calculate the required moles of H_2 for the available moles of N_2:
- For 2 moles of N_2: $2 \, \text{moles } N_2 \times \frac{3 \, \text{moles } H_2}{1 \, \text{mole } N_2} = 6 \, \text{moles } H_2$

Step 3: Since only 5 moles of H_2 are available, H_2 is the limiting reagent.

Understanding Percent Yield

Definition

Percent yield is a measure of the efficiency of a chemical reaction, expressed as the ratio of the actual yield (the amount of product obtained from the reaction) to the theoretical yield (the amount of product that could be produced based on stoichiometric calculations) multiplied by 100.

$$
\text{Percent Yield} = \left(\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \right) \times 100
$$

Importance of Percent Yield

Calculating the percent yield has several advantages:

- Quality Control: It helps in assessing the performance of a reaction in laboratory and industrial settings.
- Identifying Errors: A low percent yield can indicate issues such as incomplete reactions, side reactions, or loss of product during transfer.
- Process Optimization: Understanding the percent yield can inform adjustments to reaction conditions to improve efficiency.

Calculating Percent Yield

To calculate percent yield, follow these steps:

1. Determine the Theoretical Yield: Use stoichiometry to find the maximum amount of product that can be formed from the limiting reagent.
2. Measure the Actual Yield: This is the amount of product actually obtained from the experiment.
3. Apply the Percent Yield Formula.

Example Calculation

Continuing with the previous example of H_2 and N_2 reacting to form NH_3, assume that the actual yield of NH_3 obtained from the reaction is 4 moles.

Step 1: Calculate the theoretical yield of NH_3 based on the limiting reagent (H_2). Since 3 moles of H_2 produce 2 moles of NH_3, the theoretical yield is:
- For 5 moles of H_2:

$$
5 \, \text{moles } H_2 \times \frac{2 \, \text{moles } NH_3}{3 \, \text{moles } H_2} \approx 3.33 \, \text{moles } NH_3
$$

Step 2: The actual yield is given as 4 moles of NH_3.

Step 3: Calculate the percent yield:

$$
\text{Percent Yield} = \left(\frac{4 \, \text{moles}}{3.33 \, \text{moles}} \right) \times 100 \approx 120.12\%
$$

This yields a percent yield greater than 100%, indicating that there may have been a measurement error or a side reaction resulting in additional product.

Common Sources of Error in Percent Yield Calculations

When calculating percent yield, several factors can lead to discrepancies between theoretical and actual yields:

- Measurement Inaccuracies: Errors in measuring reactants or products can significantly affect yield calculations.

- Side Reactions: Unintended reactions can produce additional products or consume reactants, impacting the yield.
- Loss of Product: Products may be lost during transfer, purification, or isolation processes.
- Incomplete Reactions: If the reaction does not go to completion, the actual yield will be lower than expected.

Strategies to Improve Yield

Improving percent yield can be achieved through various strategies:

1. Optimize Reaction Conditions: Adjusting temperature, pressure, or concentration of reactants can enhance reaction efficiency.
2. Minimize Product Loss: Careful techniques during purification and transfer can reduce losses.
3. Use Catalysts: Catalysts can speed up reactions and help achieve higher yields.
4. Ensure Complete Reactions: Using excess reactants judiciously can drive reactions to completion.

Conclusion

In summary, understanding the concepts of limiting reagent and percent yield is essential for any chemist. These two concepts not only help in calculating the efficiency of chemical reactions but also play a crucial role in resource management and optimization in both laboratory and industrial settings. By mastering how to identify the limiting reagent and calculate percent yield, chemists can make informed decisions that lead to more effective and cost-efficient chemical processes.

Frequently Asked Questions

What is a limiting reagent in a chemical reaction?

A limiting reagent is the substance that is completely consumed in a chemical reaction, which limits the amount of product that can be formed.

How do you identify the limiting reagent in a reaction?

To identify the limiting reagent, calculate the moles of each reactant and determine which one produces the least amount of product based on the stoichiometry of the reaction.

What is percent yield in chemistry?

Percent yield is a measure of the efficiency of a reaction, calculated as the ratio of the actual yield (amount of product obtained) to the theoretical yield (maximum possible amount of product), multiplied by 100.

How do you calculate percent yield?

Percent yield is calculated using the formula: (actual yield / theoretical yield) x 100.

What happens if you have excess reactants in a reaction?

If you have excess reactants, they will remain unreacted after the limiting reagent is consumed, and they do not affect the amount of product formed.

Why is it important to know the limiting reagent in a reaction?

Knowing the limiting reagent helps predict the maximum amount of product that can be formed and allows for more efficient use of reactants in chemical processes.

Can percent yield exceed 100%?

No, percent yield cannot exceed 100%. A yield over 100% suggests that there has been an error in measurement or that impurities are present in the product.

How can you improve the percent yield of a reaction?

To improve percent yield, optimize reaction conditions such as temperature, pressure, and concentration, ensure complete reactions, and minimize side reactions and product losses.

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