If The Percent Yield Ofr The Following Reaction Is 75% And 45.0g Of NO2 Are Consumed With Ercess Water

If The Percent Yield Of The Following Reaction Is 75% And 45.0g Of NO2 Are Consumed With Excess Water, understanding the principles of chemical reactions, stoichiometry, and percent yield calculations becomes essential for accurately determining the theoretical and actual amounts of products formed. This scenario is a common problem in chemistry, especially when analyzing reaction efficiencies and yields. In this comprehensive article, we will delve into the details of this problem, exploring the underlying chemistry, calculations involved, and practical applications to help students and professionals alike understand how to approach similar chemical yield problems effectively.

Understanding the Basics: Key Concepts in Chemical Reactions and Yield

Before tackling the specific problem, it's important to understand fundamental concepts such as reaction stoichiometry, limiting reactants, theoretical yield, actual yield, and percent yield. These concepts form the foundation of quantitative chemical analysis.

What Is Stoichiometry?

Stoichiometry refers to the calculation of reactants and products involved in a chemical reaction based on the balanced chemical equation. It allows chemists to predict the amount of products formed from given reactants.

Limiting Reactant and Excess Reactant

- Limiting Reactant: The reactant that is completely consumed first and limits the amount of product formed.
- Excess Reactant: Reactant(s) that remain after the reaction has gone to completion, present in quantities more than enough to react with the limiting reactant.

Theoretical vs. Actual Yield

- Theoretical Yield: The maximum amount of product that can be formed from a given amount of reactants, based on stoichiometry.
- Actual Yield: The amount of product actually obtained from a reaction, which is often less due to various practical factors.

Percent Yield

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

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

The Chemical Reaction Involving NO_2 and Water

The specific reaction under consideration involves nitrogen dioxide (NO_2) reacting with water to produce nitric acid (HNO_3) and nitrogen monoxide (NO):

$$
2\, \text{NO}_2 + \text{H}_2\text{O} \rightarrow \text{HNO}_3 + \text{NO}
$$

This reaction is significant in environmental chemistry, industrial nitric acid production, and pollution control.

Key points about this reaction:
- NO_2 acts as an oxidizing agent.
- Water facilitates the formation of nitric acid.
- The reaction typically proceeds with excess water to ensure complete conversion of NO_2.

Given Data and Objective

The problem states:

- The percent yield is 75%.
- The mass of NO_2 consumed is 45.0 g.
- Water is in excess, meaning it does not limit the reaction.

Objective: To determine the theoretical yield of nitric acid produced and analyze the reaction efficiency based on the given data.

Step-by-Step Calculation Approach

To solve this problem systematically, follow these steps:

1. Identify the reaction and molar ratios.
2. Calculate moles of NO_2 used.
3. Determine the theoretical moles and mass of the product (HNO_3).
4. Use the percent yield to find the actual yield of HNO_3.
5. Discuss the significance and applications of the calculations.

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1. Moles of NO_2 Consumed

The molar mass of nitrogen dioxide (NO_2):

$$
\text{N} = 14.01\, \text{g/mol}
$$

$$
\text{O}_2 = 16.00\, \text{g/mol} \times 2 = 32.00\, \text{g/mol}
$$

$$
\text{Molar mass of NO}_2 = 14.01 + 32.00 = 46.01\, \text{g/mol}
$$

Calculate moles of NO_2:

$$
\text{Moles of NO}_2 = \frac{45.0\, \text{g}}{46.01\, \text{g/mol}} \approx 0.978\, \text{mol}
$$

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2. Theoretical Yield of Nitric Acid (HNO_3)

From the balanced equation:

$$
2\, \text{NO}_2 + \text{H}_2\text{O} \rightarrow \text{HNO}_3 + \text{NO}
$$

- Mole ratio: 2 mol NO_2 produce 1 mol HNO_3.

Calculate moles of HNO_3 theoretically produced:

$$
\text{Moles of HNO}_3 = \frac{1}{2} \times \text{moles of NO}_2 = \frac{1}{2} \times 0.978 \approx 0.489\, \text{mol}
$$

Molar mass of HNO_3:

$$
\text{H} = 1.008\, \text{g/mol}
$$
$$
\text{N} = 14.01\, \text{g/mol}
$$
$$
\text{O}_3 = 16.00 \times 3 = 48.00\, \text{g/mol}
$$
$$
\text{Molar mass of HNO}_3 = 1.008 + 14.01 + 48.00 = 63.02\, \text{g/mol}
$$

Calculate theoretical mass of HNO_3:

$$
\text{Theoretical mass} = 0.489\, \text{mol} \times 63.02\, \text{g/mol} \approx 30.8\, \text{g}
$$

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3. Actual Yield of HNO_3 Based on Percent Yield

Given the percent yield:

$$
75\% = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%
$$

Solve for actual yield:

$$
\text{Actual Yield} = \frac{75}{100} \times 30.8\, \text{g} \approx 23.1\, \text{g}
$$

Thus, approximately 23.1 grams of nitric acid are obtained in the reaction.

Implications and Practical Applications

Understanding how to calculate theoretical and actual yields has several practical implications:

- Industrial Production: Ensuring efficient manufacturing processes by maximizing yield.
- Environmental Chemistry: Assessing pollutant removal efficiencies in pollution control.
- Research and Development: Optimizing reaction conditions for better yields.

Key Points to Remember

- Always start with a balanced chemical equation.
- Convert given quantities to moles for accurate stoichiometric calculations.
- Use percent yield to evaluate the efficiency of a reaction.
- Excess water ensures that NO_2 is the limiting reactant, simplifying calculations.

Additional Considerations in Yield Calculations

While the above calculations provide a clear method, real-world reactions may involve additional factors:

- Reaction Completeness: Not always 100% complete.
- Side Reactions: Unintended reactions reducing the yield.
- Purity of Reactants: Impurities may affect actual yield.
- Reaction Conditions: Temperature, pressure, and catalysts influence efficiency.

Conclusion

In summary, when given the percent yield and amount of reactant consumed, it is possible to determine the theoretical yield of products in a chemical reaction. For the reaction involving NO_2 and water producing nitric acid, consuming 45.0 grams of NO_2 with a 75% yield results in an approximate actual nitric acid yield of 23.1 grams. Mastering these calculations enhances understanding of chemical efficiencies, optimizing industrial processes, and improving laboratory practices.

By understanding these principles, students and chemists can accurately assess reaction performance, troubleshoot issues, and design better chemical processes for various applications.

Frequently Asked Questions

What is the percent yield in a chemical reaction?

The percent yield is the ratio of the actual yield to the theoretical yield, expressed as a percentage, indicating how much product was obtained compared to the maximum possible amount.

How do you calculate the theoretical yield of NO from 45.0 g of NO_2?

First, write the balanced chemical equation, determine the molar ratios, calculate moles of NO_2, and then use stoichiometry to find the maximum amount of NO that can be produced, converting back to grams.

Given a 75% yield and 45.0 g of NO_2, how much NO is actually produced?

Calculate the theoretical yield of NO and then multiply by 0.75 (75%) to find the actual yield. For example, if theoretical yield is 60 g, actual yield is 45 g.

What is the significance of using excess water in this reaction?

Using excess water ensures that water is not the limiting reagent, allowing the reaction to proceed to produce the maximum amount of NO based on NO_2 consumption.

How does the percent yield affect the amount of product obtained in a reaction?

A lower percent yield means less product is obtained than theoretically possible, while a higher percent yield indicates a more efficient reaction.

What are common reasons for a reaction to have a less than 100% yield?

Common reasons include incomplete reactions, side reactions, product loss during purification, and experimental errors.

How can the percent yield be improved in a laboratory setting?

Improving reaction conditions, ensuring complete reactions, minimizing product loss, and careful handling can help increase the percent yield.

If the actual yield of NO is 45.0 g with a 75% yield, what is the theoretical yield?

Theoretical yield = actual yield / (percent yield/100) = 45.0 g / 0.75 = 60.0 g of NO.

Additional Resources

If The Percent Yield Of The Following Reaction Is 75% And 45.0g Of NO_2 Are Consumed With Excess Water

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Introduction

In the realm of chemical synthesis and industrial processes, understanding the nuances of reaction efficiency is paramount. The provided scenario—where a reaction involving nitrogen dioxide (NO_2) is executed with a 75% percent yield, and 45.0 grams of NO_2 are consumed with excess water—serves as a classic example to explore the intricacies of stoichiometry, yield calculations, and reaction dynamics. This investigation aims to dissect the problem comprehensively, elucidate underlying principles, and demonstrate practical approaches for analyzing such reactions.

Context and Significance

Nitrogen dioxide (NO_2) is a reddish-brown gas that plays a significant role in environmental chemistry, particularly in the formation of acid rain and photochemical smog. Its reactions with water are fundamental in producing nitric acid (HNO_3), a vital industrial chemical used in fertilizers, explosives, and manufacturing.

Understanding the percent yield in such reactions is critical for optimizing industrial processes, minimizing waste, and improving cost efficiency. The scenario under examination—where a reaction yields 75% of the theoretical maximum—raises important questions about reaction limitations, reagent consumption, and product formation.

The Reaction Under Consideration

The primary reaction involving NO_2 and water is as follows:

$$ 3 \mathrm{NO}_2 + \mathrm{H}_2\mathrm{O} \rightarrow 2 \mathrm{HNO}_3 + \mathrm{NO} $$

This reaction illustrates the oxidation of nitrogen dioxide in aqueous media, leading to nitric acid and nitric oxide (NO).

Key points:

- Excess water is used, so water is not the limiting reagent.
- NO_2 is the limiting reagent, with a known amount (45.0 g).
- The reaction's efficiency is 75%, indicating the actual yield is 75% of the theoretical maximum.

Step 1: Calculating the Moles of NO_2 Consumed

The first step in analyzing this process is to determine how many moles of NO_2 are involved.

Molar mass of NO_2:

- Nitrogen (N): 14.01 g/mol
- Oxygen (O): 16.00 g/mol

$$\text{Molar mass of NO}_2 = 14.01 + 2 \times 16.00 = 14.01 + 32.00 = 46.01\, \mathrm{g/mol}$$

Number of moles in 45.0 g:

$$\text{Moles of NO}_2 = \frac{45.0\, \mathrm{g}}{46.01\, \mathrm{g/mol}} \approx 0.978\, \mathrm{mol}$$

Step 2: Determining the Theoretical Yield of Nitric Acid

Given the reaction stoichiometry:

$$3\, \mathrm{mol\, NO}_2 \rightarrow 2\, \mathrm{mol\, HNO}_3$$

- For every 3 moles of NO_2, 2 moles of HNO_3 are produced.

Calculate the maximum moles of HNO_3:

$$\text{Theoretical moles of HNO}_3 = \frac{2}{3} \times 0.978 \approx 0.652\, \mathrm{mol}$$

Molar mass of HNO_3:

- N: 14.01 g/mol
- H: 1.008 g/mol
- O: 16.00 g/mol

$$\text{Molar mass of HNO}_3 = 14.01 + 1.008 + 3 \times 16.00 = 14.01 + 1.008 + 48.00 = 63.02\, \mathrm{g/mol}$$

Calculate the theoretical mass of HNO_3:

$$ \text{Mass} = 0.652\, \mathrm{mol} \times 63.02\, \mathrm{g/mol} \approx 41.07\, \mathrm{g} $$

This value represents the maximum amount of nitric acid that could be produced from 45.0 g of NO_2 under ideal conditions.

Step 3: Actual Yield of Nitric Acid

The problem states that the reaction has a 75% yield, which means:

$$ \text{Actual yield} = 75\% \times \text{Theoretical yield} $$

$$ \text{Actual mass of HNO}_3 = 0.75 \times 41.07\, \mathrm{g} \approx 30.80\, \mathrm{g} $$

Thus, approximately 30.80 grams of nitric acid are produced under the given conditions.

Deep Dive: Analyzing Reaction Efficiency and Practical Implications

The scenario highlights several vital aspects of chemical reactions, especially in industrial contexts:

- Reaction Limiting Factors: Despite using excess water, the limiting reagent is NO_2, and the percentage yield indicates inefficiencies possibly due to side reactions, incomplete conversions, or process losses.
- Reagent Consumption and Waste: The consumption of 45.0 g of NO_2 results in a significant amount of product, but understanding the actual yield helps optimize resource utilization.
- Environmental Considerations: NO_2 is a pollutant; hence, maximizing yield and minimizing waste are environmentally critical.

Factors Affecting Percent Yield

Achieving a 75% yield suggests room for optimization. Factors influencing yield include:

- Reaction Conditions: Temperature, pressure, and pH can affect reaction rates and completeness.
- Purity of Reagents: Impurities may lead to side reactions, reducing yield.
- Reaction Time: Insufficient or excessive reaction time can impact conversion.
- Separation and Purification Processes: Losses during extraction, filtration, or distillation.

Practical Applications and Industrial Relevance

Understanding the yield allows chemical engineers to:

- Estimate Raw Material Needs: For a desired amount of nitric acid, calculate how much NO_2 is required.
- Design Efficient Processes: Optimize reaction conditions to approach higher yields.
- Assess Economic Viability: Balance costs of raw materials against product output.
- Implement Environmental Controls: Minimize emissions of unreacted NO_2 or NO.

Conclusion

The analysis of if the percent yield of the following reaction is 75% and 45.0 g of NO_2 are consumed with excess water underscores the importance of stoichiometry, efficiency calculations, and process optimization. From initial molar calculations to practical implications, understanding the relationship between reagent consumption and product yield is fundamental for advancing chemical manufacturing and environmental stewardship.

In this scenario, approximately 30.80 grams of nitric acid can be expected from the given input under the specified efficiency, guiding further process adjustments to improve yield and sustainability. As chemical processes grow more complex, such detailed evaluations remain essential tools for chemists and engineers striving for efficiency and environmental responsibility.

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