

What Is The Percent Yield Of The Given Reaction If 40. G Magnesium Reacts With Excess Nitric Acid To

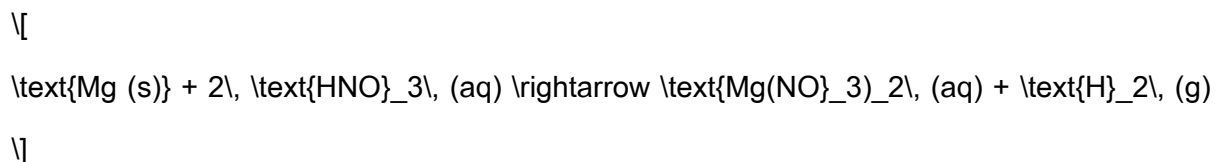
What Is The Percent Yield Of The Given Reaction If 40. G Magnesium Reacts With Excess Nitric Acid To

Understanding the concept of percent yield is fundamental in chemistry, especially when assessing the efficiency of a chemical reaction. Suppose you are given a reaction where 40.0 grams of magnesium reacts with excess nitric acid. The question arises: what is the percent yield of this reaction? To answer this, we need to analyze the reaction process, determine the theoretical yield, compare it with the actual yield, and then calculate the percent yield. This article delves into these steps in detail, providing a comprehensive understanding of how to compute the percent yield of this specific reaction.

Understanding the Reaction Between Magnesium and Nitric Acid

The Chemical Equation

The reaction between magnesium (Mg) and nitric acid (HNO_3) can be represented by the following balanced chemical equation:



This indicates that one mole of magnesium reacts with two moles of nitric acid to produce one mole of magnesium nitrate and one mole of hydrogen gas.

Reaction Details and Stoichiometry

- Reactants:
 - Magnesium (Mg): a metal, solid at room temperature
 - Nitric acid (HNO_3): a strong acid, aqueous solution, excess is assumed
- Products:
 - Magnesium nitrate ($\text{Mg}(\text{NO}_3)_2$): soluble in water
 - Hydrogen gas (H_2): released as a gas

The key point here is that magnesium is the limiting reagent, and the amount reacted determines the theoretical yield of hydrogen gas.

Calculating the Theoretical Yield

Step 1: Convert Mass of Magnesium to Moles

Given:

- Mass of magnesium = 40.0 g

Molar mass of magnesium (Mg):

$$\text{Mg} = 24.305 \text{ g/mol}$$

Number of moles of magnesium:

$$\text{Mg} =$$

$$n_{\text{Mg}} = \frac{\text{mass}}{\text{molar mass}} = \frac{40.0 \text{ g}}{24.305 \text{ g/mol}} \approx 1.646 \text{ mol}$$

\]

Step 2: Use Stoichiometry to Find Moles of Hydrogen Gas Produced

From the balanced equation:

- 1 mol Mg produces 1 mol H₂

Therefore:

\[

$$n_{\text{H}_2} = n_{\text{Mg}} \approx 1.646 \text{ mol}$$

\]

Step 3: Convert Moles of Hydrogen to Volume or Mass

Assuming standard conditions (STP), where 1 mol of gas occupies approximately 22.4 liters:

- Volume of H₂ gas:

\[

$$V_{\text{H}_2} = n_{\text{H}_2} \times 22.4 \text{ L} \approx 1.646 \times 22.4 \approx 36.9 \text{ L}$$

\]

- Mass of H₂ gas:

Molar mass of hydrogen (H₂):

\[

$$\text{H}_2 = 2.016 \text{ g/mol}$$

\]

\[

$$\text{Mass of H}_2 = n_{\text{H}_2} \times 2.016 \text{ g/mol} \approx 1.646 \times 2.016 \approx 3.32 \text{ g}$$

\]

Theoretical yield:

- The maximum amount of hydrogen that could be produced from 40 g of magnesium is approximately 3.32 g.

Understanding Actual Yield and Percent Yield

Actual Yield

The actual yield is the amount of product obtained from the reaction under real laboratory conditions. It is often less than the theoretical yield due to various factors such as incomplete reactions, side reactions, or losses during transfer.

Suppose, in an experiment, the actual amount of hydrogen gas collected is 2.5 grams.

Calculating Percent Yield

Percent yield is a measure of the efficiency of a reaction, calculated by:

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

Using the example:

$$\text{Percent Yield} = \left(\frac{2.5, \text{ g}}{3.32, \text{ g}} \right) \times 100\% \approx 75.3\%$$

This indicates that approximately 75.3% of the theoretical hydrogen was actually obtained.

Factors Affecting Percent Yield

Common Factors Influencing Reaction Efficiency

- Incomplete reactions: Not all magnesium may react
- Side reactions: Other reactions may occur that reduce hydrogen production
- Loss of product: During transfer or measurement
- Impurities: Presence of impurities in reactants
- Reaction conditions: Temperature, pressure, and concentration can influence yield

Strategies to Improve Yield

- Ensuring complete reaction by adequate stirring and contact
- Using pure reactants
- Optimizing reaction conditions
- Proper handling and minimizing losses during transfer

Practical Example: Calculating Percent Yield Step-by-Step

1. Convert the given mass of magnesium to moles: $40.0 \text{ g} \div 24.305 \text{ g/mol} \approx 1.646 \text{ mol}$

2. Determine the theoretical moles of hydrogen gas produced: same as magnesium, 1.646 mol

3. Calculate the theoretical mass of hydrogen: $1.646 \text{ mol} \times 2.016 \text{ g/mol} \approx 3.32 \text{ g}$

4. Measure the actual mass of hydrogen produced experimentally (e.g., 2.5 g)

5. Compute the percent yield: $(2.5 \text{ g} \div 3.32 \text{ g}) \times 100\% \approx 75.3\%$

Conclusion

Calculating the percent yield of a reaction like magnesium reacting with excess nitric acid involves understanding the stoichiometry of the reaction, converting reactant quantities to moles, determining the theoretical yield of the product, and then comparing it with the actual yield obtained experimentally. In the case of 40 grams of magnesium, the maximum theoretical hydrogen gas produced is approximately 3.32 grams under ideal conditions. Actual yields are often less, and the percent yield provides insight into the efficiency of the process. This calculation is essential for chemists to assess reaction performance, optimize conditions, and improve future results.

Summary of Key Points:

- Convert reactants to moles using molar mass
- Use balanced chemical equations to find mole ratios
- Calculate the theoretical yield based on limiting reagent
- Measure actual yield experimentally
- Calculate percent yield to evaluate reaction efficiency

Final note: Always consider experimental errors and practical limitations when interpreting percent yield results.

Frequently Asked Questions

What is the percent yield of the reaction if 40.0 g of magnesium reacts with excess nitric acid?

To find the percent yield, first determine the theoretical yield of the product, then compare it to the actual yield. Assuming the reaction is $\text{Mg} + 2\text{HNO}_3 \rightarrow \text{Mg}(\text{NO}_3)_2 + \text{H}_2$, the molar mass of Mg is 24.305 g/mol. Moles of Mg = $40.0 \text{ g} / 24.305 \text{ g/mol} = 1.645 \text{ mol}$. The molar ratio of Mg to $\text{Mg}(\text{NO}_3)_2$ is 1:1, so theoretical yield of $\text{Mg}(\text{NO}_3)_2 = 1.645 \text{ mol} \times \text{molar mass of } \text{Mg}(\text{NO}_3)_2 \text{ (148.32 g/mol)} = 243.9 \text{ g}$. If the actual yield is known, percent yield = $(\text{actual yield} / 243.9 \text{ g}) \times 100\%$. Without the actual yield, the percent yield cannot be calculated precisely.

How do you calculate the theoretical yield of magnesium nitrate from 40 g of magnesium reacting with nitric acid?

Calculate moles of magnesium: $40 \text{ g} / 24.305 \text{ g/mol} = 1.645 \text{ mol}$. The reaction produces 1 mol of $\text{Mg}(\text{NO}_3)_2$ per mol of Mg, so theoretical yield = $1.645 \text{ mol} \times 148.32 \text{ g/mol} = 243.9 \text{ g}$ of $\text{Mg}(\text{NO}_3)_2$.

What factors can affect the percent yield in this magnesium and nitric acid reaction?

Factors include incomplete reactions, side reactions, loss of product during transfer, measurement errors, and impurities. Proper reaction conditions and careful handling can improve yield accuracy.

If the actual amount of magnesium nitrate produced is 200 g, what is the percent yield?

Percent yield = $(200 \text{ g} / 243.9 \text{ g}) \times 100\% = 82.0\%$.

Why is it important to know the percent yield of a chemical reaction?

Percent yield indicates the efficiency of a reaction, helping chemists optimize processes, minimize waste, and assess the practicality of industrial applications.

Can the percent yield of this reaction be greater than 100%? Why or why not?

No, percent yield cannot exceed 100% because it would imply producing more product than theoretically possible, which is not physically feasible. Yields over 100% usually result from impurities or measurement errors.

What steps are involved in calculating the percent yield of magnesium nitrate in this reaction?

First, determine the theoretical yield based on stoichiometry; second, measure the actual yield obtained from the experiment; third, divide the actual yield by the theoretical yield; finally, multiply by 100% to get the percent yield.

Additional Resources

What Is The Percent Yield Of The Given Reaction If 40. G Magnesium Reacts With Excess Nitric Acid To

Understanding the percent yield of a chemical reaction is fundamental in both academic and industrial chemistry. It provides insight into the efficiency of a reaction, indicating how much of the theoretical maximum product is actually obtained. In this article, we will explore the concept of percent yield using the specific example of 40 grams of magnesium reacting with excess nitric acid. We will analyze the reaction, perform calculations to determine the theoretical and actual yields, and discuss the significance of percent yield in practical applications.

Introduction to Percent Yield in Chemical Reactions

Percent yield is a measure of the efficiency of a chemical reaction. It compares the actual amount of

product obtained from a reaction to the maximum amount possible (theoretical yield), expressed as a percentage. The formula for percent yield is:

$$\text{Percent Yield} = (\text{Actual Yield} / \text{Theoretical Yield}) \times 100\%$$

- Actual yield is the amount of product actually obtained from the reaction.
- Theoretical yield is the maximum amount of product that could be formed from the limiting reactant, assuming complete conversion and no losses.

Calculating the percent yield involves understanding stoichiometry, the molar relationships between reactants and products, and the reaction conditions.

Understanding the Reaction: Magnesium and Nitric Acid

The specific reaction in question involves magnesium reacting with nitric acid:



This is a typical single displacement reaction where magnesium displaces hydrogen from nitric acid, producing magnesium nitrate and hydrogen gas.

Key points about this reaction include:

- Magnesium is a reactive metal.
- Nitric acid is in excess, ensuring complete reaction of magnesium.
- The main product of interest is magnesium nitrate.

Stoichiometry of the Reaction

From the balanced chemical equation:

- 1 mole of Mg reacts with 2 moles of HNO_3 .
- The molar mass of magnesium (Mg) is approximately 24.305 g/mol.
- The molar mass of magnesium nitrate $[\text{Mg}(\text{NO}_3)_2]$ is approximately 148.315 g/mol.

Assuming excess nitric acid, the limiting reactant is magnesium, and the theoretical yield of magnesium nitrate can be calculated based on the initial amount of magnesium used.

Calculating Theoretical Yield

Given data:

- Mass of magnesium = 40 g

Step 1: Convert magnesium mass to moles

Number of moles of Mg:

$$n_{\text{Mg}} = \frac{\text{Mass of Mg}}{\text{Molar mass of Mg}} = \frac{40 \text{ g}}{24.305 \text{ g/mol}} \approx 1.646 \text{ mol}$$

Step 2: Determine moles of magnesium nitrate produced

From the balanced equation, the molar ratio of Mg to $\text{Mg}(\text{NO}_3)_2$ is 1:1.

Thus:

$$n_{\text{Mg(NO}_3)_2} = n_{\text{Mg}} \approx 1.646 \text{ mol}$$

Step 3: Calculate the theoretical mass of magnesium nitrate

$$\text{Mass of Mg(NO}_3)_2 = n_{\text{Mg(NO}_3)_2} \times \text{Molar mass}$$

$$\approx 1.646 \text{ mol} \times 148.315 \text{ g/mol} \approx 244.2 \text{ g}$$

Result:

The theoretical yield of magnesium nitrate is approximately 244.2 grams if all magnesium reacts completely and efficiently.

Determining Actual Yield

The actual yield is the amount of magnesium nitrate obtained experimentally. Since the user has not provided the actual yield, let's consider a hypothetical scenario where, for example, 200 grams of magnesium nitrate is obtained.

Example:

- Actual yield = 200 g

Calculating Percent Yield:

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\% = \frac{200 \text{ g}}{244.2 \text{ g}} \times 100\% \approx 81.9\%$$

This indicates that approximately 81.9% of the maximum possible magnesium nitrate was obtained in the experiment.

Factors Influencing Percent Yield

The percent yield can vary based on numerous factors:

- Reaction completeness: Not all reactants may convert to products.
- Side reactions: Other reactions may occur, reducing the amount of desired product.
- Loss during handling: Products may be lost during filtration, transfer, or purification.
- Purity of reactants: Impurities can reduce efficiency.
- Reaction conditions: Temperature, pressure, and stirring can influence reaction efficiency.

Pros and Cons of High Percent Yield

Pros:

- Maximizes resources and cost-efficiency.
- Indicates a well-optimized reaction process.
- Reduces waste and environmental impact.

Cons:

- Sometimes difficult to achieve due to practical limitations.
- Overly focusing on yield might overlook purity or safety considerations.
- High yield does not always equate to high purity of the product.

Implications of Percent Yield in Industrial and Laboratory

Settings

In industrial chemistry, high percent yields are desirable for maximizing profitability and sustainability. In laboratories, understanding yield helps in assessing reaction efficiency, optimizing procedures, and troubleshooting.

Features of yield optimization include:

- Precise control of reaction conditions.
- Using pure reactants.
- Employing appropriate purification techniques.

Potential drawbacks:

- Achieving near-100% yield can be impractical and costly.
- Sacrificing yield for purity or safety may be necessary.

Conclusion and Summary

The percent yield of a reaction provides a quantitative measure of its efficiency. In the case of magnesium reacting with excess nitric acid, the theoretical yield can be calculated based on molar ratios and initial amounts. For 40 grams of magnesium, the maximum theoretical magnesium nitrate yield is approximately 244.2 grams. The actual yield obtained in a laboratory or industrial process will determine the percent yield, reflecting the overall effectiveness of the reaction.

Understanding and calculating percent yield are essential skills in chemistry, aiding in process optimization, resource management, and quality control. Whether aiming for maximum efficiency or balancing yield with purity and safety, comprehending this concept allows chemists to design better reactions and improve outcomes.

In summary:

- Theoretical yield provides a benchmark for maximum possible product.
- Actual yield reflects real-world results, influenced by various factors.
- Percent yield is a critical metric for assessing reaction efficiency.
- Optimizing yields involves careful control of reaction conditions and handling procedures.

By mastering these concepts, chemists can better evaluate reaction success, troubleshoot issues, and enhance the productivity of chemical processes.

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