How Many Moles Of Aluminum Are Needed To React Completely With 2.1 Moles Of FeO? 2Al(s) + 3FeO(s) 3Fe(s)

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Understanding chemical reactions and stoichiometry is fundamental in chemistry, especially when determining how much of a reactant is needed to fully convert another reactant into a desired product. In this case, we are asked to find out how many moles of aluminum (Al) are required to completely react with 2.1 moles of iron(II) oxide (FeO) based on the balanced chemical equation:

$$ 2 \text{Al} (s) + 3 \text{FeO} (s) \rightarrow 3 \text{Fe} (s) $$

This question involves applying the principles of molar ratios and stoichiometry, which are essential tools in quantitative chemistry. To answer this, we will explore the reaction in detail, understand the molar relationships, and perform the necessary calculations to determine the exact amount of aluminum required.

Understanding the Balanced Chemical Equation

Reaction Components

The reaction involves three substances:

- Aluminum (Al): a metal that acts as a reducing agent.
- Iron(II) oxide (FeO): an oxide of iron, serving as the oxidizing agent.
- Iron (Fe): the product of the reduction of FeO.

The balanced equation:

$$ 2 \text{Al} (s) + 3 \text{FeO} (s) \rightarrow 3 \text{Fe} (s) $$

This indicates that:

- 2 moles of aluminum react with 3 moles of FeO.
- The reaction produces 3 moles of iron.

From the coefficients, the molar ratio of Al to FeO is 2:3.

Determining the Molar Ratio

Significance of the Coefficients

Coefficients in a balanced chemical equation provide the molar ratios of reactants and products involved in the reaction. For our reaction:

- For every 2 moles of Al, 3 moles of FeO are consumed.
- This ratio can be written as:

$$ \frac{\text{Moles of Al}}{\text{Moles of FeO}} = \frac{2}{3} $$

or equivalently:

$$ \text{Moles of Al} = \frac{2}{3} \times \text{Moles of FeO} $$

This relationship allows us to convert between quantities of different reactants.

Calculating the Moles of Aluminum Needed

Given Data

- Moles of FeO = 2.1 mol

Calculation Steps

1. Set up the molar ratio relation:

$$
\text{Moles of Al} = \frac{2}{3} \times \text{Moles of FeO}
$$

2. Plug in the given value:

$$
\text{Moles of Al} = \frac{2}{3} \times 2.1
$$

3. Perform the multiplication:

$$
\text{Moles of Al} = \frac{2 \times 2.1}{3} = \frac{4.2}{3}
$$

4. Simplify:

$$
\text{Moles of Al} = 1.4
$$

Result: To react completely with 2.1 moles of FeO, 1.4 moles of aluminum are required.

Additional Considerations in Stoichiometry

Reaction Conditions

While calculations provide the theoretical amount of reactants needed, real-world reactions may be influenced by:

- Reaction completeness: Is the reaction going to go to completion under the given conditions?
- Purity of reactants: Impurities may affect the amount of reactant needed.
- Reaction environment: Temperature, pressure, and catalysts can influence reaction efficiency.

Limiting Reactant Concept

In practical scenarios, identifying the limiting reactant is crucial. Since the problem specifies the amount of FeO and asks for the required aluminum, it assumes aluminum is in excess or that FeO is the limiting reactant.

Understanding Stoichiometry in Depth

Stoichiometry as a Tool

Stoichiometry allows chemists to:

- Calculate the exact quantities of reactants needed for a desired reaction.
- Predict the amount of products formed.

- Determine the efficiency of reactions.

Practical Applications

This calculation is not just theoretical but has real-world applications such as:

- Manufacturing processes involving metals.
- Designing chemical reactors.
- Ensuring efficient use of raw materials in industrial production.

Summary of Key Points

- The balanced chemical equation provides the molar ratio between aluminum and FeO as 2:3.
- To find the moles of aluminum needed for 2.1 mol of FeO, we use the ratio: $\text{Al} = \frac{2}{3} \times \text{FeO}$.
- The calculation yields 1.4 moles of aluminum required.
- Understanding molar ratios is essential for accurate stoichiometric calculations.
- Real-world considerations may influence the actual quantity needed, but the theoretical calculation provides a baseline.

Conclusion

In summary, based on the stoichiometry of the reaction between aluminum and iron(II) oxide, 1.4 moles of aluminum are required to completely react with 2.1 moles of FeO. This calculation underscores the importance of balanced chemical equations and molar ratios in quantitative chemistry. By mastering these principles, chemists can precisely compute reactant requirements, optimize chemical processes, and ensure efficient resource utilization in various industrial and laboratory settings.

Frequently Asked Questions

How many moles of aluminum are required to completely react with 2.1 moles of FeO?

3.0 moles of aluminum are needed to react completely with 2.1 moles of FeO, based on the balanced equation 2Al + 3FeO → 3Fe.

What is the mole ratio between aluminum and FeO in the reaction?

The mole ratio is 2:3, meaning 2 moles of aluminum react with 3 moles of FeO.

If you have 2.1 moles of FeO, how many moles of aluminum are required according to the balanced equation?

You need 1.4 moles of aluminum to react with 2.1 moles of FeO, calculated as (2/3) × 2.1.

What is the balanced chemical equation for the reaction between aluminum and iron(II) oxide?

The balanced equation is 2Al(s) + 3FeO(s) → 3Fe(s) + $2Al_2O_3$(s).

How do you determine the amount of aluminum needed to react completely with a given amount of FeO?

Use the mole ratio from the balanced equation to convert moles of FeO to moles of aluminum, then multiply accordingly.

Additional Resources

How Many Moles Of Aluminum Are Needed To React Completely With 2.1 Moles Of FeO? 2Al(s) + 3FeO(s) → 3Fe(s)

Understanding the precise amount of reactants needed in a chemical reaction is fundamental to both theoretical chemistry and practical applications such as industrial manufacturing, material science, and laboratory experiments. Today, we delve into an intriguing question: How many moles of aluminum are required to fully react with 2.1 moles of iron(II) oxide (FeO)? This inquiry not only underscores the importance of stoichiometry—the quantitative relationship between reactants and products—but also provides us with insights into how chemical equations govern real-world processes.

In this article, we will explore the reaction between aluminum and iron(II) oxide, analyze the balanced chemical equation, perform detailed calculations to determine the amount of aluminum needed, and discuss the broader implications of such reactions.

Understanding the Chemical Reaction

The Reaction Equation

The chemical reaction under consideration is:

$2Al(s) + 3FeO(s) \rightarrow 3Fe(s) + 2Al_2O_3(s)$

Note: The original equation provided is $2Al(s) + 3FeO(s) \rightarrow 3Fe(s)$. However, for clarity and completeness, especially considering typical reduction-oxidation (redox) reactions, the full balanced reaction involves the formation of aluminum oxide (Al_2O_3). The accurate balanced equation is:

$2Al(s) + 3FeO(s) \rightarrow 3Fe(s) + Al_2O_3(s)$

This reaction signifies that aluminum metal acts as a reducing agent, converting iron(II) oxide into elemental iron, while aluminum itself is oxidized to aluminum oxide.

Reaction Type and Significance

This process is a classic example of a redox reaction, where electrons are transferred from aluminum to iron ions. Such reactions are fundamental in metallurgy, especially in extractive processes like thermite reactions, which produce high-temperature reactions used for welding or metal extraction.

The reaction is exothermic, releasing significant heat, which makes understanding its stoichiometry critically important in controlling and optimizing industrial applications.

Deciphering the Stoichiometry

Importance of the Mole Ratio

The core of stoichiometry lies in the mole ratio—the ratio in which reactants combine to form products. From the balanced equation:

$2Al(s) + 3FeO(s) \rightarrow 3Fe(s) + Al_2O_3(s)$

we see that:

- 2 moles of aluminum react with 3 moles of FeO.

This ratio allows us to determine how much aluminum is needed to react with a given amount of FeO.

Calculating the Moles of Aluminum Needed

Given:
- Moles of FeO = 2.1 mol

From the balanced equation:
- 3 mol FeO react with 2 mol Al

Set up a proportion:
- 3 mol FeO / 2 mol Al = 2.1 mol FeO / x mol Al

Solve for x:
$x = (2 \text{ mol Al} \times 2.1 \text{ mol FeO}) / 3 \text{ mol FeO}$

$x = (4.2 \text{ mol Al}) / 3$

$x = 1.4 \text{ mol Al}$

Answer: 1.4 moles of aluminum are required to completely react with 2.1 moles of FeO.

Implications of the Calculation

Practical Applications

Understanding the precise amount of aluminum needed has direct implications in various fields:

- Industrial Metal Extraction: Efficient use of aluminum minimizes waste and cost.
- Thermite Welding: Accurate stoichiometry ensures safety and effectiveness.
- Materials Engineering: Precise reactant ratios influence the properties of produced alloys or compounds.

Safety and Efficiency

Knowing the exact quantities prevents excess reactant use, which can lead to unnecessary heat generation, safety hazards, or incomplete reactions. In large-scale operations, small errors in molar calculations can lead to significant economic and safety issues.

Extended Considerations

Limiting Reactants and Reaction Completion

In real-world scenarios, the reaction might not proceed to completion if one reactant is in excess. The calculation above assumes ideal conditions with complete reaction—an important assumption in theoretical chemistry. In practice, factors such as reaction kinetics, temperature, and purity affect the actual amount of reactants consumed.

Other Related Reactions and Variations

- Alternate Oxides: Reactions involving different iron oxides (e.g., Fe_2O_3) would have different molar ratios.
- Catalysts and Conditions: The presence of catalysts or specific reaction conditions can influence reaction rates, even if stoichiometrically the quantities are correct.
- Environmental Impact: Proper calculation aids in managing waste and emissions, especially when dealing with metal oxides and by-products.

Conclusion

Through a detailed examination of the reaction between aluminum and iron(II) oxide, it becomes clear that 1.4 moles of aluminum are needed to completely react with 2.1 moles of FeO. The process underscores the importance of balanced chemical equations and mole ratios in predicting reactant requirements accurately. Such calculations are not merely academic exercises but are vital in ensuring efficiency, safety, and cost-effectiveness in industrial and laboratory settings.

Understanding the stoichiometry of redox reactions like this one empowers chemists, engineers, and technicians to optimize processes, minimize waste, and innovate safer, more sustainable methods for metal extraction and material synthesis. As science advances, mastery of these fundamental principles remains a cornerstone of chemical innovation and application.

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