How Many Moles Of NH3(g) Are Produced When 12 Moles Of H2(g) Are Consumed?

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Understanding the relationship between reactants and products in chemical reactions is fundamental in chemistry. Specifically, when dealing with the synthesis of ammonia (NH_3), it's crucial to know how much product is formed from a given amount of reactants. In this article, we explore the question: **How many moles of NH_3(g) are produced when 12 moles of H_2(g) are consumed?** We will delve into the balanced chemical equation, mole ratios, and calculations to accurately determine the amount of ammonia produced.

The Haber Process: The Foundation of Ammonia Synthesis

Overview of the Haber Process

The production of ammonia primarily occurs through the Haber-Bosch process, a significant industrial method developed in the early 20th century. This process synthesizes ammonia from nitrogen and hydrogen gases under high temperature and pressure, catalyzed by iron.

The balanced chemical equation for the Haber process is:

\[
N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)
\]

This equation indicates that:

- 1 mole of nitrogen reacts with 3 moles of hydrogen.
- The reaction produces 2 moles of ammonia.

Understanding this ratio is key to calculating the amount of NH_3 formed from a specific amount of H_2.

Understanding the Mole Ratio in the Reaction

Interpreting the Balanced Equation

Based on the balanced chemical equation:

Reactant/Product	Coefficient	Interpretation
N_2	1	1 mole of nitrogen
H_2	3	3 moles of hydrogen
NH_3	2	2 moles of ammonia produced

This ratio allows us to set up conversion factors to determine the amount of ammonia produced for a given amount of hydrogen consumed.

Mole Ratio for H_2 to NH_3

From the balanced equation:

$$
3\, \text{moles H}_2 \rightarrow 2\, \text{moles NH}_3
$$

This ratio is used to calculate the moles of ammonia produced from a known amount of hydrogen.

Calculating the Moles of NH_3 Produced from 12 Moles of H_2

Step-by-Step Calculation Process

Step 1: Write the known quantities

- Moles of H_2 consumed: 12 mol

Step 2: Set up the mole ratio

Using the balanced equation:

$$
3\, \text{mol H}_2 \rightarrow 2\, \text{mol NH}_3
$$

\]

Step 3: Use proportion to find moles of NH_3

$$
\text{Moles of NH}_3 = \left(\frac{2\, \text{mol NH}_3}{3\, \text{mol H}_2} \right) \times 12\, \text{mol H}_2
$$

Step 4: Perform the calculation

$$
\text{Moles of NH}_3 = \frac{2}{3} \times 12 = 8\, \text{mol}
$$

Answer: When 12 moles of H_2 are consumed, 8 moles of NH_3 are produced.

Additional Considerations in Ammonia Production

Impact of Reaction Conditions

While the stoichiometric calculations provide the theoretical maximum amount of NH_3 produced, real-world yields may vary due to:

- Reaction efficiency
- Catalyst performance
- Temperature and pressure conditions

Optimizing these factors can lead to higher yields in industrial settings.

Limiting Reactants and Excess Reactants

In practice, nitrogen is often the limiting reactant in ammonia synthesis because hydrogen is supplied in excess. The calculations assume complete reaction of hydrogen based on the molar ratio, but actual yields depend on the availability of nitrogen.

Practical Applications of the Calculation

Understanding how to calculate the amount of ammonia produced from hydrogen consumption is essential in:

- Industrial ammonia manufacturing planning

- Designing chemical reactors
- Estimating production costs
- Environmental impact assessments

By mastering these calculations, chemists and engineers can optimize processes for maximum efficiency.

Summary of Key Points

- The balanced chemical equation for ammonia synthesis is \(N_2 + 3H_2 \rightarrow 2NH_3\).
- The molar ratio of H_2 to NH_3 is 3:2.
- Consuming 12 moles of H_2 results in the production of 8 moles of NH_3.
- Real-world yields may be affected by reaction conditions and catalyst efficiency.
- Accurate mole calculations are critical for industrial applications and process optimization.

Conclusion

To conclude, when 12 moles of H_2 gas are consumed in the synthesis of ammonia via the Haber process, 8 moles of NH_3 gas are produced, according to the stoichiometry of the balanced chemical equation. This calculation underscores the importance of understanding mole ratios in chemical reactions, enabling chemists and chemical engineers to predict product yields accurately and optimize industrial processes for maximum efficiency.

Remember: Always verify the reaction conditions and consider practical factors that may influence actual yields, but for theoretical purposes, the mole ratio provides a reliable foundation for calculations.

Frequently Asked Questions

How do I determine the moles of NH3 produced when 12 moles of H2 are consumed in the reaction?

Use the balanced chemical equation to find the molar ratio between H2 and NH3, then multiply the moles of H2 by that ratio.

What is the balanced chemical equation for the synthesis of ammonia?

The balanced equation is $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$.

If 12 moles of H2 are used, how many moles of NH3 are produced based on the reaction?

According to the balanced equation, 3 moles of H2 produce 2 moles of NH3. Therefore, 12 moles of H2 produce (12/3)×2 = 8 moles of NH3.

What is the molar ratio between H2 and NH3 in the reaction?

The molar ratio of H2 to NH3 is 3:2.

Can I directly use the molar ratio to find the amount of NH3 produced?

Yes, by multiplying the moles of H2 consumed by the ratio of NH3 to H2 from the balanced equation, you can find the moles of NH3 produced.

Why is it important to use the balanced chemical equation for these calculations?

The balanced equation ensures the correct molar ratios are used, which is essential for accurate stoichiometric calculations.

Additional Resources

How Many Moles Of NH3(g) Are Produced When 12 Moles Of H2(g) Are Consumed?

Understanding the quantitative aspects of chemical reactions is fundamental in chemistry, especially when it comes to predicting product yields based on reactant consumption. A common and illustrative example involves the synthesis of ammonia (NH_3) via the Haber-Bosch process, which is critical for fertilizer production and global agriculture. In this context, the question "How many moles of NH_3(g) are produced when 12 moles of H_2(g) are consumed?" serves as an excellent case study to explore stoichiometry, reaction mechanisms, and practical considerations in industrial chemistry.

Overview of the Haber-Bosch Process and Reaction Equation

Before diving into calculations, it's essential to understand the underlying chemical reaction. The Haber-Bosch process synthesizes ammonia from nitrogen and hydrogen gases:

$$\mathrm{N_2(g) + 3 H_2(g) \rightarrow 2 NH_3(g)}$$

This balanced chemical equation indicates that one mole of nitrogen reacts with three moles of hydrogen to produce two moles of ammonia. The reaction is exothermic and requires high temperature and pressure to proceed efficiently.

Key features of the reaction:

- Stoichiometric ratios: 1 mol N_2 : 3 mol H_2 : 2 mol NH_3
- Reaction conditions: Typically around 400–500°C and 150–200 atm
- Catalysts used: Iron-based catalysts to improve yield and rate

In our problem, nitrogen (N_2) is not specified, but since the question focuses solely on H_2 consumption and NH_3 production, we assume that nitrogen is present in excess or that the reaction proceeds with a known amount of nitrogen, and the limiting reagent is hydrogen.

Understanding the Relationship Between Reactants and Products

The key to solving this problem lies in the concept of mole ratios derived from the balanced equation. For every 3 moles of hydrogen consumed, 2 moles of ammonia are produced.

Mole ratio:

$$
\frac{\text{Moles of NH}_3}{\text{Moles of H}_2} = \frac{2}{3}
$$

This ratio is crucial because it allows us to determine the amount of ammonia produced based on the amount of hydrogen consumed.

Step-by-Step Calculation

1. Identify the Given Data

- Moles of H_2 consumed: 12 mol
- Moles of NH_3 produced: Unknown (to be calculated)
- Reaction ratio: 3 mol H_2 : 2 mol NH_3

2. Set Up the Conversion Factor

Using the mole ratio, we can set up a direct proportion:

$$
\text{Moles of NH}_3 = \left(\frac{2}{3} \right) \times \text{Moles of H}_2
$$

3. Perform the Calculation

$$
\text{Moles of NH}_3 = \left(\frac{2}{3} \right) \times 12 \text{ mol} = 8 \text{ mol}
$$

Result: When 12 moles of H_2 are consumed, 8 moles of NH_3 are produced, assuming nitrogen is in excess and the reaction goes to completion.

Implications and Assumptions in the Calculation

While the calculation appears straightforward, it’s important to recognize the assumptions made:

- Complete Reaction: The reaction proceeds to completion with no side reactions.
- Excess Nitrogen: Nitrogen is present in sufficient quantity so that hydrogen is the limiting reagent.
- Ideal Conditions: No losses, side reactions, or inefficiencies occur during the process.

In real-world industrial settings, factors such as incomplete conversions, catalyst efficiency, and operating conditions can influence the actual yield.

Factors Affecting Ammonia Production in Industrial Settings

Understanding the theoretical yield is just part of the picture. Practical considerations can significantly impact actual production quantities.

Pros of the Haber-Bosch Process:

- High Efficiency: Produces large quantities of ammonia at an industrial scale.
- Economic Viability: Cost-effective due to economies of scale.
- Recycling Capabilities: Unreacted gases can be recycled, improving overall yield.

Cons or Challenges:

- Energy Intensive: Requires high temperatures and pressures, consuming significant energy.
- Environmental Impact: The process releases greenhouse gases, especially CO_2 from energy sources.
- Catalyst Limitations: Catalyst deactivation over time can reduce efficiency.

Real-World Applications and Significance

The ability to calculate the production of ammonia from hydrogen consumption is vital for:

- Industrial Planning: Estimating raw material requirements.
- Environmental Impact Assessment: Quantifying emissions and resource use.
- Economic Analysis: Cost estimation for production scaling.

The understanding gained from stoichiometry calculations informs decisions related to process optimization, resource management, and environmental mitigation.

Additional Considerations and Extensions

While this calculation focuses solely on hydrogen, in real scenarios, nitrogen availability and purity are crucial. If nitrogen is the limiting reagent, the amount of ammonia produced would depend on the initial nitrogen amount.

Furthermore, considerations such as:

- Reaction equilibrium: Not all reactions go to completion; equilibrium

constants influence yields.
- Yield efficiency: Actual yields are often less than theoretical due to operational constraints.

These factors can be incorporated into more complex models to better predict industrial output.

Conclusion

In summary, the question “How many moles of $NH_3(g)$ are produced when 12 moles of $H_2(g)$ are consumed?” can be answered using stoichiometry directly derived from the balanced chemical equation. The calculation indicates that 8 moles of ammonia are produced under ideal, complete reaction conditions with excess nitrogen. This straightforward approach exemplifies the power of mole ratios in chemical calculations and highlights the importance of understanding reaction mechanisms, conditions, and practical considerations in industrial chemistry. Whether for academic purposes or large-scale manufacturing, mastering these concepts is essential for accurate prediction, efficient resource utilization, and sustainable chemical processes.

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