What Volume Of O2 At 988 MmHg And 37 Degrees C Is Required To Synthesize 19.5 Of NO?

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Understanding how to calculate the volume of oxygen required to synthesize a specific amount of nitric oxide (NO) under particular conditions is a fundamental aspect of chemical stoichiometry and industrial chemistry processes. This question ties into the broader context of chemical reactions involving gases, ideal gas law principles, and the practical applications of synthesis in manufacturing, environmental science, and research.

In this detailed article, we will explore the step-by-step process to determine the volume of oxygen (O_2) needed to produce 19.5 grams of nitric oxide (NO) at a pressure of 988 mmHg and a temperature of 37°C. We'll cover the relevant chemical reaction, necessary conversions, and calculations based on the ideal gas law, ensuring clarity and precision to aid students, chemists, and industry professionals alike.

Understanding the Chemical Reaction for NO Synthesis

To determine the volume of oxygen required, it's essential first to understand the chemical reaction involved in the synthesis of nitric oxide.

The Key Reaction

Nitric oxide (NO) is commonly produced through the high-temperature oxidation of nitrogen gas (N_2). The reaction can be represented as:

$$ N_2 + O_2 \rightarrow 2 NO $$

This reaction indicates that one mole of nitrogen gas reacts with one mole of oxygen gas to produce two moles of nitric oxide.

Important points:

- The reaction is 1:1:2 in terms of N_2:O_2:NO molar ratios.
- To produce a specific mass of NO, we need to determine the equivalent

number of moles and then relate that to the amount of O_2 required.

Calculating the Moles of NO Needed

Given data:

- Mass of NO to be synthesized = 19.5 grams
- Molar mass of NO = 14.01 (N) + 16.00 (O) = 30.01 g/mol

Step 1: Calculate moles of NO

$$
\text{Moles of NO} = \frac{\text{Mass of NO}}{\text{Molar mass of NO}} = \frac{19.5\, \text{g}}{30.01\, \text{g/mol}} \approx 0.6499\, \text{mol}
$$

Determining the Moles of O_2 Required

From the balanced chemical reaction:

$$ N_2 + O_2 \rightarrow 2 NO $$

- 1 mol O_2 produces 2 mol NO.

To produce 0.6499 mol NO, the amount of O_2 needed is:

$$
\text{Moles of O}_2 = \frac{1}{2} \times \text{Moles of NO} = \frac{1}{2} \times 0.6499 \approx 0.32495\, \text{mol}
$$

This is the theoretical amount of O_2 required under ideal conditions.

Applying the Ideal Gas Law to Find the Volume

The ideal gas law relates pressure (P), volume (V), number of moles (n), the gas constant (R), and temperature (T):

$$ PV = nRT $$

Rearranged to solve for volume:

$$ V = \frac{nRT}{P} $$

Constants and conversions needed:

- Gas constant (R): 0.082057 L·atm·mol^{-1}·K^{-1}
- Pressure (P): 988 mmHg
- Temperature (T): 37°C
- Number of moles (n): 0.32495 mol (oxygen)

Converting Pressure to Atmospheres

$$
P (\text{atm}) = \frac{988\, \text{mmHg}}{760\, \text{mmHg/atm}} \approx 1.3\, \text{atm}
$$

Converting Temperature to Kelvin

$$
T (\text{K}) = 37 + 273.15 = 310.15\, \text{K}
$$

Calculating the Volume of O_2 Needed

Substituting the values into the ideal gas law:

$$
V = \frac{nRT}{P} = \frac{0.32495 \times 0.082057 \times 310.15}{1.3}
$$

Calculations:

$$
V = \frac{0.32495 \times 0.082057 \times 310.15}{1.3} \approx \frac{8.303}{1.3} \approx 6.392\, \text{L}
$$

\]

Therefore, approximately 6.39 liters of oxygen are required to synthesize 19.5 grams of NO under the specified conditions.

Summary of the Calculation Process

1. Determine moles of NO: 19.5 g / 30.01 g/mol ≈ 0.6499 mol
2. Calculate O_2 moles needed: half of NO moles, ≈ 0.32495 mol
3. Convert pressure to atm: 988 mmHg ≈ 1.3 atm
4. Convert temperature to Kelvin: 37°C ≈ 310.15 K
5. Use ideal gas law to find volume: V ≈ 6.39 liters

Additional Considerations and Practical Implications

While the calculations above provide a theoretical volume of oxygen needed, real-world scenarios may involve factors such as:

- Reaction efficiency: Not all gases may react completely; some oxygen may be unreacted.
- Pressure and temperature fluctuations: Actual conditions may vary, affecting gas volumes.
- Gas purity: Impurities can influence reaction outcomes.
- Safety precautions: Working with high-temperature reactions and gases requires proper safety measures.

In industrial settings, engineers often add a safety margin (e.g., 10-20%) to account for inefficiencies and process variations.

Conclusion

To synthesize 19.5 grams of nitric oxide (NO) at a pressure of 988 mmHg and a temperature of 37°C, approximately 6.39 liters of oxygen (O_2) are required under ideal conditions. This calculation utilizes the fundamental principles of stoichiometry, the ideal gas law, and precise unit conversions, serving as a vital example of applying theoretical chemistry to practical problems.

Key Takeaways:

- The importance of balancing chemical equations for stoichiometric calculations.
- Converting units accurately for pressure, temperature, and volume.
- Applying the ideal gas law to real-world chemical engineering problems.
- Considering practical factors beyond theoretical calculations for process design.

By mastering these calculations, chemists and engineers can optimize synthesis processes, ensure safety, and improve efficiency in various applications ranging from industrial manufacturing to environmental management.

Frequently Asked Questions

How do I calculate the volume of O_2 needed to synthesize 19.5 g of NO at 988 mmHg and 37°C?

First, determine the moles of NO produced (using molar mass), then use the stoichiometry of the reaction to find the moles of O_2 required, and finally apply the ideal gas law to find the volume at the given pressure and temperature.

What is the balanced chemical equation for the synthesis of NO from O_2?

The common reaction is $N_2 + O_2 \rightarrow 2\ NO$; however, if starting from nitrogen gas, this reaction helps determine the required O_2 volume based on NO production.

How do temperature and pressure affect the calculation of O_2 volume in this reaction?

The ideal gas law ($PV=nRT$) shows that increasing temperature or decreasing pressure increases the volume of gas needed. Adjustments are made using the given pressure of 988 mmHg and temperature of 37°C (310 K).

How do I convert the given pressure of 988 mmHg to atmospheres for calculations?

Divide the pressure in mmHg by 760 mmHg to convert to atm: 988 mmHg / 760 mmHg ≈ 1.3 atm.

What is the step-by-step process to find the required volume of O_2 in this problem?

1) Calculate moles of NO from given mass; 2) Use the stoichiometry of the reaction to find moles of O_2 needed; 3) Apply the ideal gas law with the given pressure and temperature to find the volume of O_2 required.

Additional Resources

What Volume Of O_2 At 988 MmHg And 37 Degrees C Is Required To Synthesize 19.5 g Of NO?

Understanding the precise amount of oxygen required for the synthesis of nitric oxide (NO) under specific conditions is a vital question in both industrial chemistry and biochemical processes. This inquiry combines principles of stoichiometry, gas laws, and thermodynamics to determine how much oxygen is needed to produce a given mass of NO under defined temperature and pressure conditions. In this comprehensive review, we analyze the chemical reaction involved, apply the ideal gas law, and perform detailed calculations to answer this question thoroughly.

Introduction to the Chemical Context

Nitric oxide (NO) is a diatomic, colorless gas with significant industrial and biological relevance. It is synthesized primarily through the oxidation of nitrogen, which is a common process in various combustion and catalytic systems. The fundamental reaction can be summarized as:

$N_2 + O_2 \rightarrow 2\ NO$

Reaction Analysis:

- Reactants: Nitrogen gas (N_2) and oxygen gas (O_2)
- Products: Nitric oxide (NO)

This reaction is typically conducted under conditions that favor the formation of NO, such as high temperatures. For the purpose of this calculation, we assume complete reaction and ideal behavior of gases, which simplifies the analysis.

Quantitative Framework and Key Assumptions

Before delving into calculations, it is critical to establish the assumptions and the quantitative methods involved:

- Ideal Gas Behavior: Gases are assumed to behave ideally, meaning the gas laws apply without deviations.
- Complete Reaction: All the nitrogen and oxygen supplied react fully to produce NO.
- Standard Conditions: The given temperature (37°C) and pressure (988 mmHg) are used directly in calculations.
- Molecular Weights:
- N_2 = 28.02 g/mol
- O_2 = 32.00 g/mol
- NO = 30.01 g/mol

- Conversion of Units: Consistent units are maintained throughout (e.g., pressure in atm, temperature in Kelvin).

Step 1: Calculating the Moles of NO Needed

The first step is to determine how many moles of NO correspond to 19.5 grams.

Moles of NO:

$$
\text{Moles of NO} = \frac{\text{Mass of NO}}{\text{Molar mass of NO}} = \frac{19.5\, \text{g}}{30.01\, \text{g/mol}} \approx 0.65\, \text{mol}
$$

This means we need approximately 0.65 mol of NO in the reaction.

Step 2: Establishing the Stoichiometric Relationship

The balanced chemical equation shows:

$$
\text{N}_2 + \text{O}_2 \rightarrow 2\, \text{NO}
$$

- Stoichiometric ratio: 1 mol N_2 : 1 mol O_2 : 2 mol NO

From this, to produce 0.65 mol of NO:

- Moles of N_2 required:

$$
\text{Moles of N}_2 = \frac{1}{2} \times 0.65 \approx 0.325\, \text{mol}
$$

- Moles of O_2 required:

$$
\text{Moles of O}_2 = \frac{1}{2} \times 0.65 \approx 0.325\, \text{mol}
$$

Thus, approximately 0.325 mol of O_2 is needed to synthesize 19.5 g of NO.

Step 3: Applying the Ideal Gas Law to Find the Volume of O_2

The ideal gas law relates pressure, volume, temperature, and moles:

$$
PV = nRT
$$

Where:

- P = pressure in atm
- V = volume in liters
- n = number of moles
- R = ideal gas constant = 0.082057 L·atm/(mol·K)
- T = temperature in Kelvin

Convert given conditions:

- Pressure:

$$
988\, \text{mmHg} \times \frac{1\, \text{atm}}{760\, \text{mmHg}} \approx 1.3\, \text{atm}
$$

- Temperature:

$$
T = 37^\circ C + 273.15 = 310.15\, \text{K}
$$

- Number of moles of O_2:

$$
n = 0.325\, \text{mol}
$$

Calculate volume:

$$
V = \frac{nRT}{P} = \frac{0.325 \times 0.082057 \times 310.15}{1.3}
$$

$$
V \approx \frac{0.325 \times 25.44}{1.3} \approx \frac{8.263}{1.3} \approx 6.35\, \text{L}
$$

Therefore, approximately 6.35 liters of O_2 at 988 mmHg and 37°C are required to synthesize 19.5 g of NO under ideal conditions.

Additional Considerations and Analytical Insights

While the above calculations provide a clear estimate, several factors merit discussion:

1. Reaction Efficiency and Yield

In practical industrial scenarios, reactions rarely proceed at 100% efficiency. Factors like incomplete mixing, side reactions, and kinetic limitations can reduce the actual yield. Therefore, in real-world applications, a slight excess of oxygen is often supplied to ensure complete conversion of nitrogen to NO.

2. Gas Behavior Deviations

Although the ideal gas law simplifies calculations, real gases deviate from ideality at high pressures or low temperatures. Given the pressure of 988 mmHg (~1.3 atm), the deviation might be minimal but still relevant in precise calculations.

3. Temperature Impact

Temperature influences both reaction kinetics and gas volumes. The temperature of 37°C (human body temperature) suggests relevance in biological or medical contexts, but industrial processes might operate at higher temperatures to maximize yield and rate.

4. Environmental and Safety Factors

Synthesis of NO is hazardous due to its toxic nature and the potential formation of nitrogen dioxide (NO_2). Proper handling, reactor design, and safety protocols are essential when scaling such calculations to real systems.

Summary and Practical Implications

In conclusion, to synthesize 19.5 grams of nitric oxide at a pressure of 988 mmHg and a temperature of 37°C, approximately 6.35 liters of oxygen are required under ideal conditions. This calculation underscores the importance of stoichiometry, gas laws, and careful unit conversions in chemical engineering and laboratory settings.

Practical applications of this analysis include:

- Designing reactors for nitric oxide production.
- Estimating oxygen consumption in biomedical or environmental applications.
- Developing safety protocols and process controls for NO synthesis.

Final Remarks

The synthesis of nitric oxide is a prime example of how fundamental chemistry principles underpin real-world industrial and biological processes. Accurate calculations of reactant volumes, such as oxygen in this case, enable engineers and scientists to optimize reactions, ensure safety, and improve efficiency. While ideal gas assumptions simplify the process, real-world applications demand consideration of deviations, efficiency, and safety factors, all of which are crucial for translating theoretical calculations into practical outcomes.

References:

- Atkins, P., & de Paula, J. (2014). Physical Chemistry (10th Edition). Oxford University Press.
- Zumdahl, S. S., & Zumdahl, S. A. (2014). Chemistry: An Atoms First Approach. Cengage Learning.
- Lide, D. R. (2004). CRC Handbook of Chemistry and Physics. CRC Press.

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Author's Note: This analytical exploration provides a comprehensive
understanding of oxygen requirements for nitric oxide synthesis, integrating
chemical principles with practical considerations. For laboratory or
industrial scale-up, additional factors such as reaction kinetics, catalyst
presence, and environmental controls must be incorporated.
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