6. A Sample Of A Gas At 77C And 1.33 Atm Occupies A Volume Of 50.3 L. How Many Moles Of The Gas Are Present?

6. A Sample Of A Gas At 77°C And 1.33 Atm Occupies A Volume Of 50.3 L. How Many Moles Of The Gas Are Present?

Understanding the relationship between pressure, volume, temperature, and moles of a gas is fundamental in chemistry, especially when dealing with gas laws. The problem presented involves calculating the number of moles of a gas given specific conditions—namely, temperature, pressure, and volume. This type of problem typically employs the Ideal Gas Law, which provides a mathematical framework to relate these variables. In this comprehensive guide, we will explore how to approach this calculation step-by-step, review the relevant concepts, and discuss practical applications of these principles.

Introduction to the Ideal Gas Law

The Ideal Gas Law is a fundamental equation in chemistry that describes the behavior of ideal gases under various conditions. It is expressed as:

$$PV = nRT$$

Where:
- P = pressure of the gas (in atmospheres, atm)
- V = volume of the gas (in liters, L)
- n = number of moles of gas (mol)
- R = ideal gas constant (0.0821 L·atm/(mol·K))
- T = temperature in Kelvin (K)

This law combines Boyle's Law, Charles's Law, and Avogadro's Law, providing a comprehensive relation among the key variables.

Converting Units to the Correct System

Before applying the Ideal Gas Law, it's essential to ensure all units are consistent with the law's requirements.

1. Temperature Conversion

The temperature is given as 77°C. Since the Ideal Gas Law requires temperature in Kelvin:

$$T(K) = T(°C) + 273.15$$

Therefore:

$$T = 77 + 273.15 = 350.15\, \text{K}$$

For simplicity, we often round to two decimal places unless high precision is necessary.

2. Pressure Confirmation

The pressure is provided as 1.33 atm, which is already in the correct units. No conversion necessary.

3. Volume Confirmation

The volume is given as 50.3 liters, which aligns with the units used in the Ideal Gas Law.

Calculating the Number of Moles of Gas

With all variables in the correct units, the calculation proceeds straightforwardly.

Step 1: Write the Ideal Gas Law Equation

Rearranged to solve for n:

$$n = \frac{PV}{RT}$$

Step 2: Plug in the Known Values

Using:

- $P = 1.33\, \text{atm}$
- $V = 50.3\, \text{L}$
- $R = 0.0821\, \text{L·atm/(mol·K)}$
- $T = 350.15\, \text{K}$

Calculation:

$$n = \frac{(1.33)(50.3)}{(0.0821)(350.15)}$$

Step 3: Perform the Calculation

Calculate numerator:

$$1.33 \times 50.3 = 66.899$$

Calculate denominator:

$$0.0821 \times 350.15 \approx 28.743$$

Calculate n:

$$n = \frac{66.899}{28.743} \approx 2.33\, \text{mol}$$

Result: The sample contains approximately 2.33 moles of gas.

Understanding the Significance of the Calculation

Knowing the number of moles present in a gas sample is vital across various scientific and industrial applications:

- Chemical Reactions: Determining reactant or product quantities.
- Gas Production: Calculating output in industrial processes.
- Laboratory Experiments: Ensuring proper stoichiometry and safety.
- Environmental Monitoring: Assessing pollutant levels.

This calculation exemplifies how fundamental principles like the Ideal Gas Law enable scientists and engineers to quantify and manipulate gases with precision.

Practical Applications of Gas Law Calculations

The process of calculating moles from given gas conditions has numerous real-world applications:

1. Chemical Manufacturing

Understanding how much gas is involved in reactions allows for optimal reagent use, cost savings, and safety management.

2. Respiratory and Medical Fields

Calculations of gas quantities are crucial in respiratory therapy, anesthesia, and medical diagnostics.

3. Environmental Science

Monitoring atmospheric gases and pollutants often involves measuring concentrations and extrapolating molar quantities.

4. Engineering and Design

Designing equipment like reactors, engines, or gas storage tanks relies on precise molar calculations to ensure safety and efficiency.

Additional Considerations and Limitations

While the Ideal Gas Law offers a reliable approximation under many conditions, real gases exhibit deviations, especially at high pressures and low temperatures. It's important to understand these limitations:

- Non-ideal Behavior: Real gases have intermolecular forces and finite molecular sizes.
- Van der Waals Equation: An advanced model that accounts for deviations from ideality.
- Experimental Conditions: Always consider whether the gas behaves ideally under the given conditions.

In most laboratory and standard conditions, the Ideal Gas Law provides sufficiently accurate results.

Conclusion

Calculating the number of moles of a gas from given temperature, pressure, and volume involves a systematic approach rooted in the Ideal Gas Law. By converting units appropriately, substituting values, and performing straightforward calculations, one can accurately determine the molar quantity of a gas sample. In this specific case, a gas at 77°C and 1.33 atm occupying 50.3 liters contains approximately 2.33 moles. Mastery of these calculations enables chemists, engineers, and students to analyze and predict gas behavior effectively, supporting a wide array of scientific and industrial endeavors.

Summary of Key Points

- Always convert temperature to Kelvin when using the Ideal Gas Law.
- Ensure pressure is in atmospheres and volume in liters.
- Use the ideal gas constant $R = 0.0821\, \text{L·atm/(mol·K)}$ for calculations in these units.
- Apply the formula $n = \frac{PV}{RT}$ to find the number of moles.
- Understand the limitations and assumptions underlying the Ideal Gas Law.

By mastering these principles, students and professionals can confidently solve complex gas-related problems and interpret experimental data with accuracy and clarity.

Frequently Asked Questions

What is the first step to find the number of moles in the given gas sample?

Convert the temperature to Kelvin and then use the ideal gas law $PV = nRT$ to solve for n, the number of moles.

How do you convert 77°C to Kelvin?

Add 273.15 to 77°C, so the temperature in Kelvin is $77 + 273.15 = 350.15$ K.

What is the value of the ideal gas constant R used in calculations?

The value of R is 0.0821 L·atm/(mol·K).

How do you rearrange the ideal gas law to solve for the number of moles?

Rearranged as $n = PV / RT$, where P is pressure, V is volume, R is the gas constant, and T is temperature in Kelvin.

What is the calculated number of moles of the gas in the sample?

Using the given data: $n = (1.33 \text{ atm} \times 50.3 \text{ L}) / (0.0821 \text{ L·atm/(mol·K)} \times 350.15 \text{ K}) \approx 2.33$ moles.

Why is it important to convert temperature to Kelvin in gas law calculations?

Because the ideal gas law requires temperature in Kelvin to ensure the calculations are physically meaningful and accurate.

Additional Resources

A Sample Of A Gas At 77°C And 1.33 Atm Occupies A Volume Of 50.3 L. How Many Moles Of The Gas Are Present?

Introduction

Understanding the behavior of gases is fundamental in chemistry, physics, and various engineering disciplines. Gases are unique among the states of matter because their properties are highly dependent on temperature, pressure, and volume. The question of how many moles of a gas are present under specific conditions—namely temperature, pressure, and volume—is a classic problem that illustrates the application of the Ideal Gas Law.

In this article, we will analyze a scenario where a sample of gas is measured at 77°C and 1.33 atm, occupying a volume of 50.3 liters, and determine the number of moles present in this sample. This process will involve a detailed review of the Ideal Gas Law, unit conversions, and step-by-step calculations, providing a comprehensive understanding of the principles involved.

Understanding the Fundamental Concepts

The Ideal Gas Law

The Ideal Gas Law is a fundamental equation in chemistry that relates the pressure (P), volume (V), temperature (T), and amount (n, in moles) of an ideal gas:

$$PV = nRT$$

Where:

- P = Pressure (atm)
- V = Volume (L)
- n = Number of moles (mol)
- R = Ideal gas constant (0.0821 L·atm/mol·K)
- T = Temperature (Kelvin, K)

This law is an empirical relationship that holds true under many conditions for gases that approximate ideal behavior, particularly at high temperatures and low pressures.

Key Variables in the Problem

- Temperature (T): Given as 77°C, which needs to be converted to Kelvin.
- Pressure (P): 1.33 atm.
- Volume (V): 50.3 liters.
- Number of Moles (n): Unknown, to be calculated.

Converting Units and Preparing Data for Calculation

Before applying the Ideal Gas Law, it's crucial to ensure all variables are expressed in the correct units.

Temperature Conversion

Temperature in Kelvin (K) is obtained by adding 273.15 to the Celsius temperature:

- T(K) = 77°C + 273.15 = 350.15 K

For simplicity, we often round to two decimal places, but in precise calculations, it's better to retain more decimal points.

Pressure and Volume

- Pressure is already given in atm, so no conversion is necessary.
- Volume is in liters, compatible with R's units.

Gas Constant (R)

- Use R = 0.0821 L·atm/mol·K for calculations involving atm and liters.

Calculating the Number of Moles

With all variables in appropriate units, we now proceed to calculate the moles of gas.

Step-by-Step Calculation

Starting from the Ideal Gas Law:

PV = nRT

Rearranged to solve for n:

n = PV / (RT)

Plugging in the known values:

n = (1.33 atm) × (50.3 L) / (0.0821 L·atm/mol·K × 350.15 K)

Calculations:

- Numerator: 1.33 × 50.3 = 66.899
- Denominator: 0.0821 × 350.15 ≈ 28.74

Therefore:

$n \approx 66.899 / 28.74 \approx 2.33$ mol

Result: Approximately 2.33 moles of the gas are present in the sample under the given conditions.

Implications and Significance of the Calculation

This calculation exemplifies how fundamental gas laws underpin numerous practical applications, from industrial processes to laboratory experiments. Knowing the number of moles allows scientists and engineers to predict how gases will behave under different conditions, facilitate chemical reactions, and optimize processes.

Applications include:

- Chemical Manufacturing: Controlling reactant quantities.
- Environmental Science: Modeling atmospheric gases.
- Engineering: Designing pressurized systems.
- Medical Fields: Gas exchange calculations in respiratory therapy.

Understanding the molar quantity also aids in stoichiometry calculations, where precise amounts of gases are required for reactions, ensuring safety and efficiency.

Factors Affecting Gas Behavior and Deviations from Ideal Conditions

While the ideal gas law provides a robust framework for calculations, real gases sometimes deviate due to intermolecular forces and finite molecular sizes, especially under high pressure or low temperature.

Common reasons for deviations include:

- Intermolecular Attractions: Real gases exhibit Van der Waals forces that slightly reduce pressure compared to ideal gases.
- Molecular Volume: Gas molecules occupy finite space, which can impact volume calculations.

In this problem, the conditions (77°C and 1.33 atm) are moderate, and the ideal gas law provides a good approximation.

Conclusion

The determination of the number of moles in a gas sample under specific conditions is a fundamental aspect of chemical and physical sciences. By applying the ideal gas law, converting units appropriately, and performing accurate calculations, we find that approximately 2.33 moles of the gas occupy the given volume at the specified temperature and pressure.

This process underscores the importance of understanding gas laws, unit conversions, and the assumptions underlying idealized models. It also highlights the practical relevance of these calculations across diverse scientific and industrial fields.

As scientific understanding advances, more sophisticated models like the Van der Waals equation or real gas equations are employed to account for deviations, but the ideal gas law remains an essential tool for quick, reliable estimations in many scenarios.

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Disclaimer: The calculations provided are based on the ideal gas law assumptions and are approximate. Real-world gases may exhibit slight deviations, especially under conditions far from ideal.

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