A Sample Of Nitrogen Gas Expands In Volume From 1.2 To 4.0 L At A Constant Temperature. Calculate The

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The behavior of gases during expansion or compression is a fundamental aspect of thermodynamics, and nitrogen gas serves as a classic example due to its inert nature and prevalence in the atmosphere. When a sample of nitrogen gas expands from an initial volume of 1.2 liters to a final volume of 4.0 liters at constant temperature, it undergoes a process governed by the ideal gas law. Understanding the relationships between pressure, volume, and temperature during such processes allows us to calculate various properties, including the change in pressure, work done by the gas, and the amount of heat exchanged. In this article, we will explore how to perform these calculations systematically, providing a comprehensive understanding of the thermodynamics involved in the expansion of nitrogen gas at constant temperature, also known as an isothermal process.

Understanding the Fundamental Concepts

Ideal Gas Law

The ideal gas law is expressed as:

- $PV = nRT$

where:

- P = pressure of the gas
- V = volume of the gas
- n = number of moles of gas
- R = universal gas constant (8.314 J/(mol·K))

- T = absolute temperature in Kelvin (K)

Isothermal Process

An isothermal process occurs at constant temperature (T). During such a process, the internal energy of an ideal gas remains unchanged, and any work done by or on the gas is balanced by heat transfer.

Given Data and Assumptions

For this problem, the following data and assumptions are typically considered:

1. Initial volume, $V_1 = 1.2$ L
2. Final volume, $V_2 = 4.0$ L
3. The process occurs at constant temperature, T
4. The amount of nitrogen gas, n, remains constant during expansion
5. Gas behaves ideally

To proceed with calculations, we need to determine the initial pressure (P_1), the final pressure (P_2), and the work done during expansion, among other properties. For these calculations, the initial pressure P_1 must be known or assumed. If not specified, we can assume a standard initial pressure (e.g., 1 atm) for illustrative purposes.

Calculating the Initial Conditions

Determining the Number of Moles (n)

If the initial pressure P_1 is known, we can determine the number of moles using the ideal gas law:

- $n = (P_1V_1) / (RT)$

Suppose initial pressure P_1 is 1 atm (which is 101.325 kPa). Convert volume to cubic meters:

- $V_1 = 1.2\ L = 0.0012\ m^3$

Calculate n:

$n = (P_1V_1) / (RT) = (101.325 \times 10^3\ Pa \times 0.0012\ m^3) / (8.314\ J/(mol \cdot K) \times T)$

The temperature T must be specified or assumed. For typical atmospheric conditions, T ≈ 298 K (25°C).

Plugging in values:

$n = (101{,}325\ Pa \times 0.0012\ m^3) / (8.314\ J/(mol \cdot K) \times 298\ K)$
$= (121.59\ J) / (2478.57\ J/mol)$
$\approx 0.049\ mol$

Final Pressure After Expansion

Since it's an isothermal process, the relation between initial and final pressure and volume is:

- $P_1V_1 = P_2V_2$

Rearranged to find P_2:

$P_2 = (P_1V_1) / V_2$
$= (101.325\ kPa \times 1.2\ L) / 4.0\ L$
$= (101.325 \times 1.2) / 4.0$
$= 121.59 / 4.0$
$\approx 30.4\ kPa$

Calculating the Work Done During Expansion

Work in an Isothermal Process

The work done by the gas during an isothermal expansion is given by:

- $W = nRT \ln(V_2 / V_1)$

Using the previously calculated n, R, T, and the volumes:

$W = 0.049 \text{ mol} \times 8.314 \text{ J/(mol·K)} \times 298 \text{ K} \times \ln(4.0 \text{ L} / 1.2 \text{ L})$
$= 0.049 \times 8.314 \times 298 \times \ln(3.333)$

Calculate the natural logarithm:

$\ln(3.333) \approx 1.203$

Now, compute W:

$W \approx 0.049 \times 8.314 \times 298 \times 1.203$
$= 0.049 \times 8.314 \times 298 \times 1.203$
$\approx 0.049 \times 8.314 \times 358.5$
$\approx 0.049 \times 2984.1$
$\approx 146.2 \text{ J}$

This is the work done by the gas during expansion.

Heat Transfer During the Process

Applying First Law of Thermodynamics

The first law states:

- $\Delta U = Q - W$

In an ideal gas, the internal energy U depends only on temperature, and since T remains constant, $\Delta U = 0$. Therefore, the heat exchanged Q is equal to the work done:

$Q = W \approx 146.2 \text{ J}$

This indicates that approximately 146.2 Joules of heat are absorbed by the nitrogen gas during the expansion process to maintain constant temperature.

Summary of Key Calculations

- Initial moles of nitrogen: approximately 0.049 mol (assuming initial pressure of 1 atm)
- Final pressure after expansion: approximately 30.4 kPa
- Work done during expansion: approximately 146.2 Joules
- Heat absorbed during expansion: approximately 146.2 Joules

Additional Considerations and Variations

Effect of Different Initial Pressures

If the initial pressure P_1 differs from 1 atm, the calculations for n, P_2, and work would need adjustment accordingly. The general approach remains the same: use the ideal gas law to find n, then apply the relations for isothermal processes.

Non-ideal Gas Behavior

At high pressures or low temperatures, nitrogen may deviate from ideal behavior. In such cases, real gas equations like the Van der Waals equation provide more accurate predictions, but for typical laboratory or atmospheric conditions, the ideal gas law suffices.

Other Thermodynamic Processes

In practice, processes may not be perfectly isothermal. To analyze adiabatic or isobaric expansions, different formulas and considerations are necessary, often involving entropy and specific heat capacities.

Conclusion

The expansion of nitrogen gas at constant temperature illustrates key principles of thermodynamics and gas laws. By applying the ideal gas law and understanding the relationships between pressure, volume, and temperature, we can accurately calculate the work done, the change in pressure, and the heat transfer involved. Such calculations are fundamental in fields ranging from chemical engineering to atmospheric science, providing insights into the behavior of gases under various conditions. Whether for academic purposes or practical applications, mastering these principles enables a deeper comprehension of the energetic transformations that gases undergo during expansion or compression.

Frequently Asked Questions

A sample of nitrogen gas expands from 1.2 L to 4.0 L at constant temperature. What is the ratio of the initial and final volumes?

The ratio of initial to final volume is 1.2 L to 4.0 L, which simplifies to 1.2 / 4.0 = 0.3.

Given the expansion of nitrogen gas at constant temperature, what law can be used to relate the initial and final volumes?

Boyle's Law can be used, which states that for a fixed amount of gas at constant temperature, the pressure and volume are inversely proportional (PV = constant).

If the initial pressure of nitrogen gas is 1 atm, what is the final pressure after expansion, assuming the temperature remains constant?

Using Boyle's Law, P1V1 = P2V2. Therefore, P2 = P1V1 / V2 = 1 atm 1.2 L / 4.0 L = 0.3 atm.

How does the temperature of nitrogen gas change during the volume expansion at constant temperature?

The temperature remains constant during the expansion, as specified in the problem statement, due to isothermal conditions.

What is the significance of constant temperature in calculating the expansion of nitrogen gas?

Constant temperature allows the use of Boyle's Law, simplifying calculations by keeping the product of pressure and volume constant.

Calculate the work done by the nitrogen gas during its expansion from 1.2 L to 4.0 L at constant temperature, assuming initial pressure is 1 atm.

Work done (W) = nRT ln(V2/V1). Without the number of moles (n), pressure, or temperature specified, the exact work cannot be calculated; additional data is needed.

What assumptions are made when calculating gas expansion at constant temperature and volume change?

Assumptions include ideal gas behavior, no heat exchange with surroundings (isothermal process), and that the gas does not experience any phase change or chemical reaction.

Additional Resources

Nitrogen Gas Expansion at Constant Temperature: A Comprehensive Analysis and Calculation

Understanding the behavior of gases under varying conditions is fundamental in chemistry and physics. When a gas expands or compresses, the change in its volume, pressure, and temperature can be described and predicted using the fundamental gas laws. In this detailed exploration, we will analyze a specific scenario involving nitrogen gas expanding from 1.2 liters to 4.0 liters at constant temperature and perform the relevant calculations to understand what is happening at the molecular level.

Introduction to Gas Behavior and the Ideal Gas Law

Before delving into the specific problem, it is essential to review the principles governing gases. Gases are composed of molecules in constant, random motion, and their macroscopic properties—pressure (P), volume (V), temperature (T), and amount of substance (n)—are interconnected via the ideal gas law:

$$ PV = nRT $$

Where:

- P = pressure of the gas (in atm, Pa, etc.)
- V = volume (in liters, cubic meters, etc.)

- n = number of moles
- R = ideal gas constant (~8.314 J/(mol·K))
- T = temperature (in Kelvin)

Since the problem specifies that temperature remains constant, the scenario can be directly analyzed using Boyle's Law.

Boyle's Law: The Foundation for Isothermal Processes

Boyle's Law states that, for a fixed amount of gas at constant temperature:

$$ P_1V_1 = P_2V_2 $$

This implies that pressure and volume are inversely proportional under isothermal conditions. When a gas expands (volume increases), the pressure decreases proportionally, assuming temperature and molar amount remain unchanged.

Key points:

- Constant Temperature (Isothermal Process): No change in the internal energy of an ideal gas.
- Inverse Relationship: As volume increases, pressure decreases proportionally.
- Applicability: Ideal gases at moderate pressures and temperatures.

In our case, nitrogen gas expands from 1.2 L to 4.0 L at the same temperature, so the pressure change can be derived using Boyle's Law.

The Scenario: Nitrogen Gas Expansion from 1.2 L to 4.0 L

Let's outline the given data:

- Initial volume, $V_1 = 1.2\, \text{L}$
- Final volume, $V_2 = 4.0\, \text{L}$
- Temperature, T (constant)
- The initial pressure, P_1, is unknown but can be expressed in terms of known quantities.
- The number of moles, n, remains unchanged because the process is

assumed to be closed with no addition or removal of gas.

The primary goal here is to calculate the change in pressure (or other related parameters, if specified). Since the problem's title indicates "Calculate the," but does not specify what, we'll explore possible calculations, including pressure change, work done during expansion, and the implications for the system.

Calculating the Pressure Change Using Boyle's Law

Given the initial and final volumes at a constant temperature, the pressure change can be directly calculated if one of the pressures is known. If the initial pressure, P_1, is provided, then:

$$P_2 = \frac{P_1 V_1}{V_2}$$

Assuming an initial pressure: Let's assume the initial pressure, P_1, is 1 atm for simplicity.

- Initial conditions:
$$P_1 = 1\, \text{atm}$$
$$V_1 = 1.2\, \text{L}$$

- Final conditions:
$$V_2 = 4.0\, \text{L}$$
$$P_2 = ?$$

Applying Boyle's Law:

$$P_2 = \frac{P_1 V_1}{V_2} = \frac{1\, \text{atm} \times 1.2\, \text{L}}{4.0\, \text{L}} = 0.3\, \text{atm}$$

This indicates that as nitrogen expands, the pressure drops from 1 atm to 0.3 atm.

Implications:

- Pressure decrease: The gas experiences a significant reduction in pressure during expansion.
- Work done by the gas: The gas performs work on its surroundings during expansion, which we will explore further in the next section.

Work Done During Isothermal Expansion

In thermodynamics, the work performed by an ideal gas during an isothermal expansion or compression is given by:

$$ W = nRT \ln \frac{V_2}{V_1} $$

Where:

- W = work done (Joules)
- n = number of moles
- R = ideal gas constant (8.314 J/(mol·K))
- T = temperature in Kelvin
- V_1, V_2 = initial and final volumes

Calculating the work:

1. Determine the number of moles, n:

Using the ideal gas law at initial conditions:

$$
n = \frac{P_1 V_1}{RT}
$$

Assuming the initial pressure $P_1 = 1\, \text{atm}$ and converting units:

- $1\, \text{atm} = 101.325\, \text{kPa}$
- $V_1 = 1.2\, \text{L} = 0.0012\, \text{m}^3$
- $R = 8.314\, \text{J/(mol·K)}$
- Temperature T (unknown)—let's assume a typical laboratory temperature, say 298 K (25°C).

Calculating n:

$$
n = \frac{(101.325 \times 10^{3}\, \text{Pa}) \times (0.0012\, \text{m}^3)}{8.314\, \text{J/(mol·K)} \times 298\, \text{K}}
$$

$$
n \approx \frac{121.59\, \text{J}}{2477.57\, \text{J/mol}} \approx 0.0491\, \text{mol}
$$

2. Calculate work done:

$$
W = nRT \ln \frac{V_2}{V_1}
$$

\]

$$
W = 0.0491\, \text{mol} \times 8.314\, \text{J/(mol·K)} \times 298\, \text{K} \times \ln \left(\frac{4.0}{1.2} \right)
$$

Calculate each component:

- $nRT = 0.0491 \times 8.314 \times 298 \approx 121.59\, \text{J}$ (consistent with previous calculation)
- $\ln \left(\frac{4.0}{1.2} \right) = \ln (3.333) \approx 1.203$

Thus,

$$
W \approx 121.59\, \text{J} \times 1.203 \approx 146.3\, \text{J}
$$

Interpretation:

- The gas performs approximately 146.3 Joules of work during this expansion.
- Since the process is isothermal, the internal energy change is zero; the work done is balanced by heat exchange with the surroundings.

Heat Exchange During Isothermal Expansion

In an ideal gas undergoing an isothermal process, the internal energy change (ΔU) is zero because internal energy depends solely on temperature:

$$
\Delta U = 0
$$

From the First Law of Thermodynamics:

$$
\Delta U = Q - W
$$

Where:

- Q = heat absorbed by the system
- W = work done by the system

Given $\Delta U = 0$, then:

$$
Q = W \approx 146.3\, \text{J}
$$

This indicates that approximately 146.3 Joules of heat energy are absorbed from the surroundings to facilitate the expansion.

Implications for Molecular Behavior and Practical Applications

Understanding the microscopic implications of this expansion provides insight into molecular dynamics:

- Molecular Motion: As the volume increases, molecules have more space to move, reducing collision frequency and pressure.
- Energy Transfer: Heat absorption compensates for the work done, maintaining constant temperature.
- Real-World Applications:

- Industrial Gas Processes: Understanding expansion behaviors is crucial in designing turbines, engines, and gas storage systems.
- Environmental Impact: Nitrogen expansion is relevant in atmospheric phenomena and in the design of pneumatic systems.
- Scientific Research: Precise measurements of gas laws help calibrate instruments and validate thermodynamic theories.

Additional Considerations and Real-World Deviations

While the ideal gas law provides a good approximation, real gases exhibit deviations under certain conditions:

- High Pressure or Low Temperature: Intermolecular

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