One Mole Of An Ideal Gas Is Expanded From A Volume Of 1.00 Liter To A Volume Of 8.93 Liters Against A

One Mole Of An Ideal Gas Is Expanded From A Volume Of 1.00 Liter To A Volume Of 8.93 Liters Against A scenario presents an intriguing case for understanding the fundamental principles of thermodynamics, particularly the behavior of ideal gases during expansion processes. This article delves into the detailed analysis of such an expansion, exploring the concepts of work done, heat transfer, and the thermodynamic equations that govern these processes. Whether you're a student preparing for exams or a science enthusiast interested in the mechanics of gases, this comprehensive overview will clarify the essential principles involved.

Understanding the Basics of Ideal Gas Expansion

What Is an Ideal Gas?

An ideal gas is a theoretical gas composed of point particles that do not interact with each other except during elastic collisions. The behavior of ideal gases is described by the Ideal Gas Law:

$$ PV = nRT $$

where:
- $ P $ = pressure,
- $ V $ = volume,
- $ n $ = number of moles,
- $ R $ = universal gas constant,
- $ T $ = temperature in Kelvin.

This law provides a foundation for analyzing gas behaviors under various conditions.

Expansion Processes of Gases

Expansion of gases can be categorized based on how the process occurs:

- Isothermal Expansion: Temperature remains constant ($ T = \text{constant} $)
- Adiabatic Expansion: No heat exchange with surroundings ($ Q = 0 $)
- Isobaric Expansion: Pressure remains constant
- Isochoric Process: Volume remains constant

In this case, since the problem involves a change in volume, understanding the type of expansion process is critical.

Analyzing the Expansion from 1.00 Liter to 8.93 Liters

Given Data

- Number of moles, $n = 1$ mol
- Initial volume, $V_i = 1.00$ L
- Final volume, $V_f = 8.93$ L
- Gas constant, $R = 8.314\, \text{J mol}^{-1} \text{K}^{-1}$

The key question is: Against what external pressure or force is the gas expanding? Since the statement is incomplete, typical scenarios include expansion against a constant external pressure, against a vacuum, or in a controlled environment. For the purpose of this article, we will analyze the expansion against a constant external pressure, which is a common assumption in thermodynamics.

Assumptions for the Analysis

- The gas behaves ideally throughout the expansion.
- The process occurs at a constant temperature (isothermal process).
- The external pressure is constant during the expansion.

These assumptions simplify calculations and allow us to illustrate fundamental principles clearly.

Calculating Work Done During Expansion

Work in Thermodynamics

Work (W) done by a gas during expansion or compression is given by:

$$W = \int_{V_i}^{V_f} P_{\text{ext}}\, dV$$

where P_{ext} is the external pressure.

- For constant external pressure, the work simplifies to:

$$W = P_{\text{ext}} (V_f - V_i)$$

- For an ideal gas undergoing an isothermal process, the internal pressure varies, but assuming a quasi-static process, the work can be calculated as:

$$W = nRT \ln \frac{V_f}{V_i}$$

since for an isothermal process, the temperature remains constant, and from the ideal gas law, $P = \frac{nRT}{V}$.

Case 1: Expansion Against a Constant External Pressure

Suppose the external pressure P_{ext} is known. For example, if the external pressure is equal to the initial pressure of the gas:

$$P_i = \frac{nRT}{V_i}$$

then the work done is:

$$W = P_i (V_f - V_i)$$

Case 2: Isothermal Expansion

If the expansion is isothermal at temperature T, then the work done by the gas is:

$$W = nRT \ln \frac{V_f}{V_i}$$

This formula is derived from integrating the pressure over volume during an isothermal process.

Example Calculation:

Let's assume the gas is at a constant temperature of 300 K.

Calculate the work done:

$$W = nRT \ln \frac{V_f}{V_i} = (1\, \text{mol})(8.314\, \text{J mol}^{-1} \text{K})(300\, \text{K}) \ln \frac{8.93}{1.00}$$

$$W = 8.314 \times 300 \times \ln(8.93)$$

$$W \approx 2494.2 \times 2.188$$

$$W \approx 5452\, \text{J}$$

Thus, approximately 5452 Joules of work are done by the gas during this expansion under isothermal conditions.

Heat Transfer During Expansion

First Law of Thermodynamics

The first law states:

$$\Delta U = Q - W$$

where:
- ΔU = change in internal energy,
- Q = heat added to the system,
- W = work done by the system.

If the expansion is isothermal:

- $\Delta U = 0$ (since internal energy of an ideal gas depends only on temperature)
- Therefore, $Q = W$

This indicates that the heat absorbed by the gas during expansion equals the work done by the gas.

If the expansion is adiabatic:

- $Q = 0$, so:

$$\Delta U = - W$$

- The internal energy decreases as the gas expands, and no heat is exchanged with the surroundings.

Implications and Real-World Applications

Practical Applications of Gas Expansion

Understanding gas expansion is essential in numerous industrial and natural processes, including:

- Internal Combustion Engines: where gases expand rapidly, doing work to move pistons.
- Refrigeration Cycles: involving isothermal and adiabatic expansions and compressions.
- Aerospace Engineering: analyzing the behavior of gases in propulsion systems.
- Climate Science: understanding atmospheric gas behaviors and thermodynamic cycles.

Designing Systems for Controlled Expansion

Engineers utilize the principles outlined above to design systems that maximize efficiency:

- Controlling Pressure and Temperature: to optimize work output or energy conservation.
- Using Isothermal or Adiabatic Processes: depending on the application's thermal management needs.
- Calculating Work and Heat Transfer: for energy budgeting and system optimization.

Conclusion

Analyzing the expansion of one mole of an ideal gas from 1.00 liter to 8.93 liters involves understanding key thermodynamic principles, including the ideal gas law, work done during expansion, and heat transfer mechanisms. Whether considering an isothermal process at constant temperature or an adiabatic process with no heat exchange, the calculations hinge on fundamental equations that describe the behavior of gases under changing conditions. Recognizing the nature of the external forces acting against the gas and understanding how to apply the relevant thermodynamic equations enables scientists and engineers to predict and harness the energy transformations involved in gas expansions effectively.

In summary:

- The work done during expansion can be calculated using the ideal gas law and the specific expansion process assumptions.
- Heat transfer depends on whether the process is isothermal, adiabatic, or occurs under other conditions.
- Such analyses are vital in designing engines, turbines, refrigeration systems, and understanding natural phenomena.

For further exploration, consider varying the temperature or external pressure conditions, or analyzing real gases that deviate from ideal behavior at high pressures or low temperatures.

Frequently Asked Questions

What is the work done during the expansion of one mole of an ideal gas from 1.00 L to 8.93 L against a constant external pressure?

The work done is given by $W = P_{external} \times \Delta V$. If the external pressure is known, multiply it by the change in volume (8.93 L - 1.00 L) to find the work done in joules or appropriate units.

How does the type of process (isothermal, adiabatic, etc.) affect the work done during this expansion?

The process type determines how temperature and pressure change during expansion. For an ideal gas, isothermal expansion involves work $W = nRT \ln(V_{final}/V_{initial})$, while adiabatic expansion involves different relations. The specific process dictates the calculation of work done.

What is the significance of the initial and final volumes in calculating the thermodynamic work of expansion?

The initial and final volumes determine the change in volume (ΔV), which directly influences the amount of work performed during expansion, especially in processes where work is proportional to volume change.

If the expansion is against a constant external pressure, how do you determine whether the work done is positive or negative?

Work done by the system during expansion is considered positive because the system does work on the surroundings. Conversely, if compression occurs, the work would be negative.

How does the ideal gas law relate to the expansion process from 1.00 L to 8.93 L?

The ideal gas law PV = nRT helps relate pressure, volume, and temperature during the process. If temperature remains constant (isothermal), the pressure decreases proportionally with volume; if not, temperature and pressure change according to the process specifics.

What additional information is needed to fully analyze the thermodynamics of this expansion process?

Details such as the external pressure, temperature conditions, whether the process is isothermal or adiabatic, and whether the process is reversible or irreversible are necessary to accurately calculate work, heat transfer, and changes in internal energy.

Additional Resources

Ideal Gas Expansion Analysis: A Deep Dive into Thermodynamic Behavior

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Introduction

Understanding the behavior of gases during expansion processes is fundamental in thermodynamics, especially when analyzing real-world applications such as engines, refrigeration systems, and chemical reactors. Today, we dissect a classic yet insightful scenario: "One mole of an ideal gas is expanded from a volume of 1.00 liter to 8.93 liters against a... (pressure, which we will specify later)." This case provides a comprehensive window into the principles governing ideal gases, including work done during expansion, changes in internal energy, heat transfer, and entropy. Our goal is to offer an in-depth, expert perspective on this process, exploring all relevant thermodynamic quantities and their implications.

The Scenario in Focus

Imagine a controlled environment where one mole of an ideal gas initially occupies a volume of 1.00 liter. The gas then undergoes an expansion to 8.93 liters. The process occurs against a specified external pressure, which critically influences the thermodynamic pathways and the energy exchanges involved.

While the original prompt does not specify the external pressure (A), for comprehensive analysis, we will consider different scenarios:

- Isobaric expansion: The process occurs at constant external pressure.
- Adiabatic expansion: No heat exchange with surroundings.
- Isothermal expansion: The temperature remains constant during the process.

In real applications, the nature of expansion—whether it's free, isothermal, adiabatic, or otherwise—dictates the energy transfer mechanisms and the work performed by or on the gas. For clarity and depth, we will analyze each case, emphasizing their thermodynamic distinctions.

Fundamental Thermodynamic Principles

Before delving into specific cases, it’s crucial to establish the core principles that underpin the analysis:

- Ideal Gas Law:
$$ PV = nRT $$
where P is pressure, V is volume, n is the number of moles, R is the universal gas constant, and T is temperature in Kelvin.

- Work Done by the Gas:
$$ W = \int P_{\text{ext}} \, dV $$
The work depends on the external pressure and the volume change.

- Change in Internal Energy:
$$ \Delta U = n C_V \Delta T $$
For an ideal gas, internal energy depends solely on temperature.

- Heat Transfer:
$$ Q = \Delta U + W $$
From the first law of thermodynamics.

- Entropy Change:
$$ \Delta S = n R \ln \frac{V_f}{V_i} \quad \text{(for isothermal)} $$

Analyzing the Expansion: Step-by-Step Breakdown

1. Initial Conditions and Constants

- Moles of gas, $n = 1$ mol.
- Initial volume, $V_i = 1.00$ L.
- Final volume, $V_f = 8.93$ L.
- Gas constant, $R = 8.314 \, \text{J/(mol·K)}$.

The initial pressure and temperature are not specified, but for a typical analysis, assume an initial temperature T_i. We could choose a standard temperature, for instance, 273 K (0°C), or analyze

the process generally.

Case 1: Isothermal Expansion

Definition: The temperature remains constant throughout the process ($T = T_i = T_f$).

Implications:
- Since T is constant, the internal energy change $\Delta U = 0$ for an ideal gas.
- The work done by the gas during expansion is directly related to the change in volume.

Calculations:

- Initial Pressure:
$$P_i = \frac{nRT}{V_i}$$
$$P_i = \frac{(1)(8.314)(273)}{1 \times 10^{-3} \, \text{m}^3} \approx 2.27 \times 10^6 \, \text{Pa}$$
(Note: 1 L = 1 x 10^-3 m^3)

- Work Done:
For isothermal expansion against an external pressure that matches the internal pressure (quasi-static process), the work is:
$$W = nRT \ln \frac{V_f}{V_i}$$
$$W = (1)(8.314)(273) \ln(8.93)$$
$$W \approx 8.314 \times 273 \times 2.19 \approx 4,971 \, \text{J}$$

- Heat Transfer:
$$Q = W$$
Since $\Delta U = 0$, all heat supplied to the system appears as work done during expansion.

- Entropy Change:
$$\Delta S = n R \ln \frac{V_f}{V_i}$$
$$\Delta S = 1 \times 8.314 \times 2.19 \approx 18.2 \, \text{J/K}$$

Expert Insight:
This process exemplifies ideal isothermal expansion—a process where the gas does maximum work while maintaining a constant temperature. The significant increase in volume results in a notable entropy increase, indicating greater disorder. However, it requires continuous heat input to maintain the temperature, reflecting an energy transfer from the surroundings.

Case 2: Adiabatic Expansion

Definition: No heat exchange occurs ($Q = 0$).

Implications:
- The process is internally conservative, with energy changes solely due to work.
- Temperature decreases as the gas expands.

Calculations:

- Adiabatic Relations:
$$ TV^{\gamma-1} = \text{constant} $$
$$ P V^\gamma = \text{constant} $$
where $\gamma = \frac{C_P}{C_V}$. For a diatomic ideal gas (e.g., N_2, O_2), $\gamma \approx 1.4$.

- Final Temperature:
$$ T_f = T_i \left(\frac{V_i}{V_f} \right)^{\gamma - 1} $$
$$ T_f = 273 \times \left(\frac{1}{8.93} \right)^{0.4} $$
$$ T_f \approx 273 \times (0.112)^{0.4} $$
$$ T_f \approx 273 \times 0.370 \approx 101 \, \text{K} $$

- Work Done:
$$ W = \frac{n R (T_i - T_f)}{\gamma - 1} $$
$$ W = \frac{1 \times 8.314 (273 - 101)}{0.4} $$
$$ W \approx \frac{8.314 \times 172}{0.4} \approx \frac{1,431}{0.4} \approx 3,577 \, \text{J} $$

- Entropy Change:
Since the process is not isothermal, the entropy change for the system is:
$$ \Delta S = n C_V \ln \frac{T_f}{T_i} + n R \ln \frac{V_f}{V_i} $$

For diatomic gases, $C_V = \frac{R}{\gamma - 1} \approx \frac{8.314}{0.4} = 20.785 \, \text{J/(mol·K)}$.

$$ \Delta S = 1 \times 20.785 \times \ln \frac{101}{273} + 1 \times 8.314 \times 2.19 $$
$$ \approx 20.785 \times (-0.996) + 18.2 $$
$$ \approx -20.7 + 18.2 = -2.5 \, \text{J/K} $$

Expert Insight:
The negative entropy change indicates that the system cools during adiabatic expansion, decreasing its entropy. However, since the universe's total entropy increases (via entropy transfer to surroundings), the overall entropy of the universe still increases, satisfying the second law. The process inherently involves energy conversion from internal energy to work, with no heat exchange.

Case 3: Isobaric Expansion with External Pressure A

Definition: The external pressure P_A remains constant during the expansion.

Suppose A is a specific pressure value—say, $P_A = 1 \text{ atm} = 101.3 \, \text{kPa}$.

Calculation of Work:

- Work Done:
$$ W = P_A (V_f - V_i) $$
$$ W = 101.3 \

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