When The Temperature Of 0.0788 Moles Of A Diatomic Gas Drops, Its Internal Energy Drops By -50.7 J. What

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Understanding the relationship between temperature and internal energy in gases is fundamental in thermodynamics. When analyzing a scenario where a specific amount of a diatomic gas experiences a temperature drop, resulting in a change in its internal energy, it becomes essential to interpret the physical implications and the underlying principles. This article explores the thermodynamic concepts involved, calculations to determine temperature change, and the significance of the data provided.

Introduction to Internal Energy and Diatomic Gases

What Is Internal Energy?

Internal energy (U) of a system refers to the total energy contained within the system, including kinetic and potential energies at the molecular level. For gases, especially ideal gases, the internal energy predominantly depends on the kinetic energy of molecules, which is directly related to temperature.

Properties of Diatomic Gases

Diatomic gases, such as oxygen (O_2), nitrogen (N_2), and hydrogen (H_2), consist of molecules with two atoms. These gases have specific degrees of freedom that influence their internal energy and heat capacity:

- Translational: 3 degrees of freedom
- Rotational: 2 degrees of freedom (for non-linear molecules, but diatomic molecules are linear)
- Vibrational: 1 or more modes (at higher temperatures)

At moderate temperatures, vibrational modes are often not excited, and the molar heat capacity remains relatively constant.

Thermodynamics Principles Relevant to the Scenario

Internal Energy Change and Temperature

For an ideal diatomic gas, the change in internal energy (ΔU) is directly proportional to the change in temperature (ΔT):

$$
\Delta U = n C_v \Delta T
$$

Where:
- n = number of moles of gas
- C_v = molar heat capacity at constant volume
- ΔT = change in temperature

Molar Heat Capacity for Diatomic Gases

For diatomic gases at room temperature (assuming vibrational modes are inactive):

$$
C_v = \frac{5}{2} R
$$

Where:
- R = universal gas constant $= 8.314\, \text{J/(mol·K)}$

This arises because diatomic gases have 5 degrees of freedom contributing to their internal energy (3 translational + 2 rotational).

Calculating the Temperature Drop

Given Data:
- Moles of gas, $n = 0.0788\, \text{mol}$
- Change in internal energy, $\Delta U = -50.7\, \text{J}$

From the relation:

$$
\Delta U = n C_v \Delta T
$$

Rearranged to find ΔT:

$$
\Delta T = \frac{\Delta U}{n C_v}
$$

Substituting known values:

$$
C_v = \frac{5}{2} R = \frac{5}{2} \times 8.314\, \text{J/(mol·K)} = 20.785\, \text{J/(mol·K)}
$$

Now,

$$
\Delta T = \frac{-50.7\, \text{J}}{0.0788\, \text{mol} \times 20.785\, \text{J/(mol·K)}}
$$

Calculating denominator:

$$
0.0788 \times 20.785 \approx 1.637\, \text{J/K}
$$

Finally,

$$
\Delta T \approx \frac{-50.7}{1.637} \approx -30.96\, \text{K}
$$

Conclusion: The temperature of the diatomic gas drops by approximately 31 K.

Implications of the Temperature Drop

Understanding the Magnitude of Change

A temperature decrease of approximately 31 Kelvin indicates a significant cooling event at the molecular level. Such a drop affects the kinetic energy of the molecules, leading to a decrease in pressure if volume remains constant or an expansion if the system is allowed to expand.

Relevance in Real-World Applications

- Industrial processes: Managing cooling in gas systems.

- Thermodynamic cycles: Understanding energy transfers.
- Environmental science: Studying atmospheric gases and temperature fluctuations.

Additional Considerations and Assumptions

Assumption of Ideal Gas Behavior

The calculations assume that the diatomic gas behaves ideally, meaning interactions between molecules are negligible, and the internal energy depends solely on temperature. Real gases may deviate slightly, especially at high pressures or low temperatures.

Vibrational Modes and Temperature Range

At higher temperatures, vibrational modes become excited, increasing the heat capacity. If vibrational modes are active, C_v would be higher, leading to a different temperature change for the same internal energy variation.

Isothermal or Adiabatic Processes

The calculation presumes a process where the internal energy change is solely due to temperature change, without additional work or heat transfer considerations. In practical settings, external factors could influence the process.

Summary of Key Points

- The internal energy of a diatomic ideal gas is proportional to temperature, with $C_v = \frac{5}{2} R$.
- A drop of 50.7 J in internal energy for 0.0788 mol of diatomic gas corresponds to approximately a 31 K temperature decrease.
- The calculation relies on the assumption of ideal gas behavior and vibrational modes being inactive at the given temperature range.
- Understanding these relationships is crucial for thermodynamics, engineering, and scientific research involving gases.

Conclusion

In conclusion, when the temperature of 0.0788 moles of a diatomic gas drops, causing its internal energy to decrease by 50.7 J, the temperature decreases by about 31 Kelvin. This insight stems from fundamental thermodynamic principles relating internal energy, heat capacity, and temperature change. Recognizing these relationships allows scientists and engineers to predict system behaviors, optimize processes, and understand the energy dynamics within gaseous systems.

Further Reading and Resources

- Thermodynamics Textbooks: For in-depth understanding of internal energy and heat capacities.
- Engineering Thermodynamics: Practical applications involving gas behavior.
- Online Calculators: Tools for quick thermodynamic calculations.
- Research Articles: Recent studies on diatomic gases and energy transfer processes.

By mastering these concepts, one can better analyze and predict the behavior of gases under various thermal conditions.

Frequently Asked Questions

What is the relationship between the internal energy change and temperature change for a diatomic gas?

For a diatomic gas, the internal energy change (ΔU) is related to temperature change (ΔT) by $\Delta U = n C_v \Delta T$, where C_v is the molar heat capacity at constant volume, typically 5/2 R. Therefore, a change in internal energy corresponds directly to a change in temperature.

Given the internal energy drop of -50.7 J for 0.0788 moles of a diatomic gas, how can we determine the change in temperature?

Using the relation $\Delta U = n C_v \Delta T$, rearranged as $\Delta T = \Delta U / (n C_v)$. For a diatomic gas, $C_v = 5/2 R \approx 20.78$ J/(mol·K). Substituting the values: $\Delta T = -50.7 \text{ J} / (0.0788 \text{ mol} \times 20.78 \text{ J/(mol·K)}) \approx -30.7$ K.

What does the negative sign in the internal energy change indicate about the temperature?

The negative sign indicates that the internal energy decreases, which corresponds to a decrease in temperature of approximately 30.7 K for the gas.

How does this energy change relate to the processes occurring in the gas?

A decrease in internal energy suggests the gas is losing heat or doing work on its surroundings, leading to a temperature drop. In this case, it indicates cooling of the gas.

What assumptions are made in calculating the temperature change of the diatomic gas?

The calculation assumes the gas behaves ideally, the process occurs at constant volume (so only internal energy changes), and the molar heat capacity C_v remains constant over the temperature range.

If the temperature of the gas drops by approximately 30.7 K, what is the initial temperature if the temperature after the drop is 250 K?

The initial temperature would be approximately 250 K + 30.7 K = 280.7 K, assuming the temperature decreased by 30.7 K.

Additional Resources

When The Temperature Of 0.0788 Moles Of A Diatomic Gas Drops, Its Internal Energy Drops By -50.7 J: An Investigative Analysis

Introduction

Understanding the behavior of gases under varying thermal conditions is fundamental to thermodynamics and physical chemistry. When analyzing the internal energy changes of gases, especially diatomic gases, the relationship between temperature and internal energy becomes a crucial point of examination. This article delves into a specific scenario: what occurs when the temperature of 0.0788 moles of a diatomic gas decreases sufficiently to result in an internal energy loss of 50.7 joules? We aim to explore the underlying principles, calculations, and implications of this phenomenon, providing a comprehensive review suitable for researchers, educators, and students alike.

Background: Internal Energy and Diatomic Gases

Internal energy (U) of an ideal gas is primarily a function of temperature, reflecting the sum of microscopic kinetic and potential energies of particles. For ideal gases, potential energy contributions are negligible, and internal energy depends solely on temperature.

In diatomic gases such as nitrogen (N_2), oxygen (O_2), or hydrogen (H_2), the internal energy accounts for:

- Translational motion (movement of molecules along x, y, z axes)
- Rotational motion (spinning of molecules)

- Vibrational motion (oscillation of atoms within molecules)

The extent to which vibrational modes contribute depends on temperature; at lower temperatures, vibrational modes are often "frozen out" due to quantum effects.

Theoretical Framework

The change in internal energy (ΔU) for an ideal gas is related to the change in temperature (ΔT) by:

$\Delta U = n C_v \Delta T$

Where:

- n = number of moles
- C_v = molar heat capacity at constant volume
- ΔT = change in temperature in Kelvin (K)

For diatomic gases, the molar heat capacity at constant volume (C_v) varies with temperature:

- At low temperatures (vibrational modes inactive): $C_v \approx (5/2) R$
- At higher temperatures (vibrational modes active): C_v approaches $(7/2) R$

Here, R is the universal gas constant (8.314 J/(mol·K)).

Given the data, the scenario suggests an internal energy decrease of 50.7 J when the temperature drops, and the number of moles is specified as 0.0788 mol. Our goal is to determine the change in temperature and analyze the implications.

Calculations and Derivations

Step 1: Determine molar heat capacity (C_v)

Given the internal energy change (ΔU) and the number of moles (n), we can write:

$\Delta U = n C_v \Delta T$

Rearranged to find ΔT:

$\Delta T = \Delta U / (n C_v)$

However, since C_v depends on temperature (and vibrational modes), we need to consider the appropriate value of C_v for the temperature range in question.

Step 2: Assumptions about the gas and temperature

Given the small change in internal energy, it’s reasonable to assume the temperature change is moderate enough that vibrational modes remain inactive, and $C_v \approx (5/2) R = 2.5\ 8.314 \approx 20.785$ J/(mol·K).

Using this, we can estimate the temperature change:

ΔT = -50.7 J / (0.0788 mol 20.785 J/(mol·K))
= -50.7 J / (1.637 J/K)
≈ -30.96 K

This indicates that the temperature drops by approximately 31 Kelvin.

Step 3: Validation and implications

If the initial temperature ($T_{initial}$) was, for example, 300 K, then:

T_{final} ≈ 300 K - 31 K = 269 K

This temperature decrease is consistent with typical thermodynamic processes involving cooling of gases.

Deeper Analysis: Quantum Considerations and Vibrational Modes

While the above calculations assume classical equipartition, quantum effects at lower temperatures may influence the heat capacity. Vibrational modes in diatomic gases typically activate at higher temperatures, usually above 500 K, depending on the molecule.

- If vibrational modes are inactive, C_v ≈ (5/2) R
- If vibrational modes are partially active, C_v increases towards (7/2) R, or approximately 29.1 J/(mol·K)

Suppose vibrational modes are partially excited. Then, C_v could be around 25 J/(mol·K). Repeating the calculation:

ΔT = -50.7 J / (0.0788 mol 25 J/(mol·K))
≈ -50.7 J / 1.97 J/K
≈ -25.7 K

This suggests a slightly larger temperature drop for the same internal energy change.

Implications for Thermodynamic Processes

This analysis highlights several key points:

- The internal energy change of a diatomic gas is directly tied to the temperature change, scaled by the number of moles and heat capacity.
- Small internal energy drops correspond to modest temperature decreases, which can be significant in precise thermodynamic control or experimental setups.
- Quantum effects can alter heat capacities, especially at low temperatures, influencing the energy-temperature relationship.

Practical Considerations and Applications

Understanding these relationships is vital in fields such as:

- Atmospheric science, where temperature variations affect gas behaviors
- Industrial processes involving gas cooling or heating
- Design of thermodynamic cycles, including refrigeration and engine efficiency optimization
- Fundamental research in quantum thermodynamics and molecular physics

Experimental Methods for Verification

To verify such a scenario experimentally, one would employ:

- Calorimetry: measuring heat exchange during temperature changes
- Spectroscopic techniques: analyzing vibrational mode activation
- Precise temperature sensors and pressure measurements to track thermodynamic parameters

Conclusion

The investigation into the cooling of 0.0788 moles of a diatomic gas, leading to an internal energy decrease of 50.7 J, underscores the delicate interplay between molecular motion, quantum effects, and thermodynamic principles. The estimated temperature drop of around 31 Kelvin, assuming classical heat capacities, exemplifies how small energetic changes can have measurable thermal consequences. Recognizing the role of vibrational mode activation and quantum considerations refines our understanding of thermodynamic behavior in diatomic gases, with broad implications across scientific and engineering disciplines.

Future research directions include exploring non-ideal gas effects, transient heat transfer phenomena, and the influence of external fields on internal energy changes. Such studies will deepen our comprehension of molecular energetics and enhance the precision of thermodynamic models in complex systems.

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