Under Constant-volume Conditions, 2700 J Of Heat Is Added To 1.5 Moles Of An Ideal Gas. As A Result,

Under Constant-volume Conditions, 2700 J Of Heat Is Added To 1.5 Moles Of An Ideal Gas. As A Result, several fundamental thermodynamic processes occur that illustrate the principles of heat transfer, internal energy change, and work done within the system. Understanding what happens during such a scenario is essential for grasping key concepts in thermodynamics, especially when analyzing ideal gases under controlled conditions. This article explores the implications of adding heat to an ideal gas at constant volume, including the calculations involved, the physical interpretations, and the broader applications in scientific and engineering contexts.

Understanding the Basic Principles of Heat Addition at Constant Volume

What is a Constant-volume Process?

A constant-volume process refers to a thermodynamic transformation where the volume of the system remains unchanged throughout the process. In practical terms, this is often achieved by confining the gas within a rigid container that does not expand or contract during heating or cooling. Since the volume is fixed, any heat added or removed from the system directly influences the internal energy and temperature of the gas, without performing work on the surroundings due to volume change.

Significance of Adding Heat to an Ideal Gas

When heat is added to an ideal gas at constant volume, the primary effects include an increase in the internal energy and temperature of the gas. Because ideal gases follow the kinetic theory, the temperature is directly proportional to the average kinetic energy of the molecules. Therefore, any increase in heat input results in faster-moving molecules and elevated internal energy, which can be quantified through thermodynamic equations.

Calculations and Quantitative Analysis of Heat Addition

Given Data

- Heat added, $Q = 2700$ J
- Number of moles, $n = 1.5$ mol
- Process: Constant volume

Determining the Change in Internal Energy (ΔU)

For an ideal gas, the first law of thermodynamics states:

$$\Delta U = Q - W$$

Where:
- ΔU is the change in internal energy
- Q is the heat added to the system
- W is the work done by the system

Since the process occurs at constant volume, the work done by the gas is zero because work in thermodynamics is given by:

$$W = P\Delta V$$

And $\Delta V = 0$, hence:

$$W = 0$$

Therefore:

$$\Delta U = Q = 2700 \text{ J}$$

This means the internal energy of the gas increases by 2700 J as a direct result of the heat added.

Calculating the Temperature Change

For an ideal gas, the internal energy change 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 for an ideal gas

For a monatomic ideal gas, $C_v = (3/2) R$
For a diatomic ideal gas (e.g., N_2, O_2), $C_v \approx R$
Where R is the universal gas constant, approximately 8.314 J/(mol·K)

Assuming the gas is diatomic (common for many gases at room temperature), then:

$$C_v = R = 8.314 \text{ J/(mol·K)}$$

Rearranging to find ΔT:

$$\Delta T = \Delta U / (n C_v) = 2700 \text{ J} / (1.5 \text{ mol} \times 8.314 \text{ J/(mol·K)}) \approx 216.2 \text{ K}$$

This indicates that the gas's temperature increases by approximately 216.2 Kelvin.

Physical and Practical Implications of Heat Addition at Constant Volume

Increase in Internal Energy and Temperature

Adding heat results in a proportional increase in the internal energy and temperature of the gas. Since no work is performed (volume remains fixed), all the heat energy contributes to raising the energy state of the molecules, leading to higher molecular velocities and kinetic energy.

Pressure Changes in the System

According to the ideal gas law:

$$PV = nRT$$

Given that the volume (V) and number of moles (n) are constant, the pressure (P) is directly proportional to the temperature (T). Therefore, as the temperature increases by approximately 216.2 K, the pressure within the container also increases proportionally, assuming the container can withstand the pressure.

Work Done During the Process

Since the volume is constant, no work is performed by the gas during heating:

$$W = P\Delta V = 0$$

This simplifies the energy transfer analysis, focusing on heat and internal energy.

Broader Applications and Significance

Engineering and Thermodynamic Cycles

Understanding heat addition at constant volume is crucial for designing engines, refrigerators, and other thermodynamic systems. For instance, the combustion chamber in an engine often operates under constant-volume conditions during combustion, making this analysis directly applicable.

Laboratory and Experimental Settings

In controlled experiments, scientists frequently heat gases in rigid containers to study temperature-dependent properties or to simulate conditions in stars or planetary atmospheres.

Refrigeration and Heating Systems

Insights into how gases respond to heat at fixed volumes aid in optimizing cooling and heating systems, ensuring safety and efficiency.

Additional Considerations and Assumptions

Ideal Gas Approximation

The calculations assume the gas behaves ideally, meaning interactions between molecules are negligible, and the molecules themselves occupy insignificant volume. Real gases deviate from ideal behavior at high pressures or low temperatures, which should be considered in practical applications.

Specific Heat Capacity Variations

The value of C_v varies with temperature and molecular complexity. For precise calculations, especially at high temperatures, the specific heat capacities should be adjusted accordingly.

Container Strength and Safety

As temperature and pressure increase, the physical container must withstand these changes. Safety considerations are essential when heating gases in confined spaces.

Summary

In conclusion, adding 2700 J of heat to 1.5 moles of an ideal gas at constant volume results in a significant increase in the internal energy (~2700 J) and temperature (~216.2 K). Since no work is done due to volume constancy, all the heat energy directly translates into higher molecular kinetic energy and pressure within the system. This fundamental process highlights the core principles of thermodynamics and provides a basis for understanding more complex systems in engineering, physics, and chemistry. Whether in designing engines, conducting laboratory experiments, or analyzing natural phenomena, the principles outlined here serve as a foundational concept for energy transfer in gases.

Frequently Asked Questions

What is the effect of adding 2700 J of heat to 1.5 moles of an ideal gas under constant-volume conditions?

The internal energy of the gas increases, leading to a rise in temperature, since at constant volume all the heat added goes into increasing the internal energy.

How much is the change in internal energy for the gas

when 2700 J of heat is added at constant volume?

The change in internal energy is approximately 2700 J, since for an ideal gas at constant volume, $\Delta U = Q$, assuming no work is done.

What is the temperature change of the gas after adding 2700 J of heat?

Using $\Delta U = n C_v \Delta T$, the temperature change can be calculated. For an ideal gas, $\Delta T = Q / (n C_v)$.

If the molar heat capacity at constant volume (C_v) for the gas is known, how can we find the temperature increase?

Temperature increase $\Delta T = Q / (n C_v)$. For example, if C_v is in J/(mol·K), divide 2700 J by (1.5 mol × C_v) to find ΔT.

Does any work get done during this process under constant-volume conditions?

No, since the volume remains constant, no work is done ($W = 0$). All the heat added increases the internal energy.

Is the process an example of an isochoric process?

Yes, because the volume remains constant, making this an isochoric (constant-volume) process.

Can we determine the final temperature of the gas after heat addition?

Yes, by using the relation $\Delta U = n C_v \Delta T$ and knowing the initial temperature, or by calculating ΔT from the heat added, if the initial temperature is known.

What assumptions are made about the gas in this problem?

The gas is assumed to be ideal, and the process is assumed to be at constant volume with no additional work done or heat losses.

How does the heat capacity at constant volume influence the temperature change?

A higher C_v means a smaller temperature increase for the same amount of heat added, as ΔT is inversely proportional to C_v.

Additional Resources

Thermodynamics: Exploring the Impact of Constant-Volume Heating on an Ideal Gas

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Introduction

Understanding how gases respond to heat under different conditions is fundamental in thermodynamics, a branch of physics with applications spanning from engineering to atmospheric science. In this detailed review, we analyze a specific scenario: adding 2700 Joules of heat to 1.5 moles of an ideal gas at constant volume. This situation offers a rich insight into the principles governing energy transfer, work, and temperature change in gases—a topic that has both theoretical significance and practical implications.

Fundamentals of Thermodynamics in Gases

Before delving into the specifics of the scenario, it's essential to establish the foundational concepts of thermodynamics applicable here:

- First Law of Thermodynamics: $\Delta U = Q - W$

- ΔU: Change in internal energy

- Q: Heat added to the system

- W: Work done by the system

- Ideal Gas Law: $PV = nRT$

- P: Pressure

- V: Volume

- n: Number of moles

- R: Universal gas constant

- T: Absolute temperature

- Internal Energy of an Ideal Gas: $U = \frac{f}{2} nRT$

- f: Degrees of freedom (for monatomic gases, $f=3$; for diatomic gases, typically $f=5$ at room temperature)

Scenario Breakdown: Heating at Constant Volume

In this particular case, the key process is constant-volume heating. This means:

- The volume V remains unchanged throughout the process.

- The heat $Q = 2700\, \mathrm{J}$ is added to the gas.

- The number of moles $n = 1.5$.

Understanding the implications of constant volume is crucial because:

- No work is done by the gas during the process ($W=0$).

- All the heat added contributes solely to changing the internal energy and, consequently, the temperature.

Implication of Constant Volume on Work and Internal Energy

Since the volume remains constant:

- Work Done (W): $W = P \Delta V = 0$, because $\Delta V = 0$.

- Change in Internal Energy (ΔU): $\Delta U = Q$, as per the first law, simplifies to:

$$
\Delta U = Q - W = Q
$$

This directly links the heat added to the internal energy increase, making calculations straightforward.

Calculating the Increase in Internal Energy and Temperature

Given the heat added, the internal energy change can be computed using the thermodynamic relations for ideal gases.

Step 1: Determine the Internal Energy Change (ΔU)

$$
\Delta U = Q = 2700\, \mathrm{J}
$$

Because the system is ideal, the internal energy depends solely on temperature, and the change in internal energy is directly proportional to the temperature change.

Step 2: Find the Internal Energy Per Mole

For a monatomic ideal gas (such as helium, neon, or argon), the internal energy per mole is:

$$
U_{per\, mol} = \frac{3}{2} RT
$$

For diatomic gases (like nitrogen or oxygen), at room temperature, the internal energy per mole is:

$$
U_{per\, mol} = \frac{5}{2} RT
$$

The specific heat capacity per mole at constant volume (C_v) for these gases is:

- Monatomic: $C_v = \frac{3}{2} R$

- Diatomic: $C_v = \frac{5}{2} R$

Given the problem statement doesn't specify the gas, we'll analyze both cases:

Case 1: Monatomic Ideal Gas

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

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

Where:

$$
C_v = \frac{3}{2} R \approx \frac{3}{2} \times 8.314\, \mathrm{J/(mol\,K)} \approx 12.471\, \mathrm{J/(mol\,K)}
$$

Plugging in the values:

$$
\Delta T = \frac{2700\, \mathrm{J}}{1.5\, \mathrm{mol} \times 12.471\, \mathrm{J/(mol\,K)}} \approx \frac{2700}{18.7065} \approx 144.2\, \mathrm{K}
$$

Result: The temperature increases by approximately 144.2 K.

Case 2: Diatomic Ideal Gas

$$
C_v = \frac{5}{2} R \approx 20.785\, \mathrm{J/(mol\,K)}
$$

Similarly:

$$
\Delta T = \frac{2700}{1.5 \times 20.785} \approx \frac{2700}{31.177} \approx 86.6\, \mathrm{K}
$$

Result: The temperature increases by approximately 86.6 K.

Analyzing the Results: Temperature and Internal Energy

The calculations show that adding heat at constant volume results in:

- A proportional increase in temperature: the greater the heat added, the higher the temperature rise.

- An increase in internal energy: since $\Delta U = Q$, the internal energy increases by 2700 J.

This relationship underscores a fundamental principle: at constant volume, all heat added goes into increasing the internal energy and temperature—no work is performed since the volume remains unchanged.

Additional Considerations and Real-World Applications

While the calculations are straightforward for ideal gases, real-world gases often deviate from ideal behavior, especially at high pressures or low temperatures. Nonetheless, the principles remain instructive.

Practical applications include:

- Calorimetry: Measuring specific heat capacities by observing temperature changes upon heating.

- Engine design: Understanding how gases respond to heat under different constraints.

- Material processing: Precise control of temperature in manufacturing processes involving gases.

Limitations and Assumptions

- Ideal gas assumption: Real gases may exhibit interactions that alter internal energy and behavior.

- Constant molar quantity: The number of moles remains fixed; mass loss or gain isn't considered.

- No phase change: The gas stays in the gaseous phase; no condensation or other phase transitions occur.

- Uniform heating: Heat distribution is even throughout the gas, avoiding temperature gradients.

Summary of Key Findings

- Heat added: 2700 J at constant volume leads to an increase in internal energy of the same magnitude.

- Temperature increase:

- Monatomic gas: approximately 144.2 K

- Diatomic gas: approximately 86.6 K

- Internal energy change: directly proportional to the heat added, with no work done during the process.

- Process implications: The process is characterized by an increase in temperature and internal energy, with no volume change or work output.

Concluding Remarks

This analysis highlights the elegant simplicity of thermodynamic principles when applied to ideal gases. The scenario of adding heat at constant volume vividly demonstrates how energy transfer translates into temperature and internal energy changes, providing essential insights for students, engineers, and scientists alike. Understanding these fundamental concepts equips practitioners with the tools needed to manipulate and predict gas behavior in diverse applications, from laboratory experiments to industrial processes.

Whether designing heat engines or conducting calorimetry, recognizing the relationship between heat, internal energy, and temperature underpins much of thermodynamics—a testament to the enduring importance of these principles in science and engineering.

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