Calculate The Value Of H When Temp Of 1 Mole Of A Monoatomic Gas Increased From 25 To 300 C Deltau

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Understanding the thermodynamic properties of gases is essential in fields ranging from chemical engineering to physical chemistry. When dealing with a monoatomic gas, calculating the enthalpy change (ΔH) during a temperature increase involves comprehending the relationship between internal energy, heat capacities, and temperature. This article provides a comprehensive guide to calculating the enthalpy change (H) when 1 mole of a monoatomic gas undergoes a temperature increase from 25°C to 300°C, focusing on the value of ΔH in this context.

Fundamentals of Thermodynamics for Monoatomic Gases

Before diving into calculations, it’s crucial to understand the basic thermodynamic principles that govern monoatomic gases.

What is a Monoatomic Gas?

- A monoatomic gas consists of single-atom particles, such as helium (He), neon (Ne), or argon (Ar).
- These gases have simple atomic structures, leading to straightforward thermodynamic properties.
- Their behavior can often be accurately modeled using the ideal gas law and classical thermodynamics.

Key Thermodynamic Quantities

- **Internal Energy (U):** The total energy contained within a system, primarily kinetic energy for monoatomic gases.
- **Enthalpy (H):** Defined as H = U + PV, where P is pressure and V is volume.
- **Heat Capacity at Constant Volume (Cv):** The amount of heat required to raise the

temperature of the gas by one degree at constant volume.

- **Heat Capacity at Constant Pressure (Cp):** The heat needed to raise the temperature by one degree at constant pressure.

Thermodynamic Relationships for Monoatomic Gases

The next step involves understanding the relationships between these quantities for monoatomic gases.

Heat Capacities

- For monoatomic ideal gases, the molar heat capacity at constant volume (Cv) is given by:

$Cv = (3/2) R$

- Where R is the universal gas constant, approximately 8.314 J/(mol·K).
- The molar heat capacity at constant pressure (Cp) is related to Cv by:

$Cp = Cv + R = (3/2) R + R = (5/2) R$

Internal Energy and Enthalpy

- The change in internal energy (ΔU) for a monoatomic ideal gas during a temperature change is:

$\Delta U = n\ Cv\ \Delta T$

- Similarly, the change in enthalpy (ΔH) is:

$\Delta H = n\ Cp\ \Delta T$

Calculating ΔH for a Monoatomic Gas from 25°C to 300°C

Now, let's focus on calculating the enthalpy change (ΔH) for 1 mole of a monoatomic gas when its temperature increases from 25°C to 300°C.

Step 1: Convert Temperatures to Kelvin

- Temperature in Kelvin is obtained by adding 273.15 to Celsius temperature.

$T_1 = 25°C + 273.15 = 298.15$ K

$T_2 = 300°C + 273.15 = 573.15$ K

Step 2: Calculate ΔT

- The temperature difference is:

$\Delta T = T_2 - T_1 = 573.15\ K - 298.15\ K = 275\ K$

Step 3: Apply the Formula for ΔH

- Given that $\Delta H = n\ C_p\ \Delta T$, and for 1 mole of a monoatomic ideal gas, $C_p = (5/2)\ R$.

$\Delta H = 1\ mol \times (5/2)\ R \times \Delta T$

- Substitute $R = 8.314$ J/(mol·K):

$\Delta H = (5/2) \times 8.314 \times 275$

- Calculate the numerical value:

$\Delta H = 2.5 \times 8.314 \times 275 \approx 2.5 \times 2284.85 \approx 5712.13$ J

Summary of Calculation

The enthalpy change (ΔH) when 1 mole of a monoatomic ideal gas increases in temperature from 25°C to 300°C is approximately **5712.13 Joules**.

Additional Considerations

Although the above calculation is straightforward for ideal gases, real gases may deviate due to interactions and non-ideal behavior. Here's what to keep in mind:

Real Gas Effects

- At high pressures or low temperatures, gases may deviate from ideal behavior.
- In such cases, correction factors or real gas equations (like Van der Waals equation) should be used.

Non-Monoatomic Gases

- For diatomic or polyatomic gases, heat capacities differ, and the calculations involve different Cp values.
- For example, diatomic gases like oxygen (O_2) have $Cv \approx (5/2)\,R$ and $Cp \approx (7/2)\,R$ at room temperature.

Practical Applications of ΔH Calculations

Calculating the enthalpy change is vital in various practical scenarios:

1. **Design of Heating Systems:** Determining energy requirements for heating gases in industrial processes.
2. **Thermodynamic Efficiency:** Assessing the energy changes in engines and turbines.
3. **Chemical Reactions:** Understanding heat absorption or release during reactions involving gases.

Conclusion

In summary, calculating the enthalpy change (ΔH) when 1 mole of a monoatomic gas increases from

25°C to 300°C involves understanding the relationship between heat capacity and temperature change. Using the specific heat capacity at constant pressure (Cp) for monoatomic gases, the calculation becomes straightforward. For the given temperature range, the approximate ΔH is about 5712 Joules, providing valuable insights into energy requirements and thermodynamic behavior of monoatomic gases.

Frequently Asked Questions

How do you calculate the change in enthalpy (ΔH) when the temperature of 1 mole of a monoatomic gas increases from 25°C to 300°C?

For a monoatomic ideal gas, ΔH can be calculated using $\Delta H = n C_p \Delta T$, where $C_p = (5/2) R$. Convert temperatures to Kelvin, find ΔT, then multiply by n and C_p to find ΔH.

What is the value of ΔH for 1 mole of a monoatomic gas when the temperature increases from 25°C to 300°C?

Using $\Delta H = n C_p \Delta T$ with $C_p = (5/2) R$, ΔT = 300°C - 25°C = 275°C = 275 K. Therefore, ΔH = 1 mol × (5/2) R × 275 K ≈ 1 mol × 20.785 J/K × 275 K ≈ 5714 J.

Why is it valid to assume monoatomic gases have a specific heat capacity C_p = (5/2) R for this calculation?

Monoatomic gases have only translational degrees of freedom, leading to a molar heat capacity at constant pressure of $C_p = (5/2) R$, which is a standard approximation based on kinetic theory.

How does the change in internal energy (ΔU) relate to the change in enthalpy (ΔH) for an ideal gas during heating?

For an ideal gas, $\Delta U = n C_v \Delta T$ and $\Delta H = n C_p \Delta T$. Since $C_p = C_v + R$, the difference between ΔH and ΔU is R T, but at constant pressure, ΔH and ΔU are related through their respective heat capacities.

What is the significance of calculating ΔH when heating a monoatomic gas from 25°C to 300°C?

Calculating ΔH helps determine the heat energy absorbed or released during heating, which is essential in thermodynamics for process analysis, energy management, and understanding the behavior of gases under temperature changes.

Additional Resources

Calculate The Value Of H When Temp Of 1 Mole Of A Monoatomic Gas Increased From 25 To 300°C ΔU

Understanding the thermodynamics of gases, especially monoatomic gases, involves delving into the fundamental concepts of internal energy (U), enthalpy (H), and how these properties change with temperature. In this detailed review, we will explore how to calculate the change in enthalpy (ΔH) for one mole of a monoatomic gas when its temperature increases from 25°C to 300°C, with a focus on the underlying principles, relevant equations, and step-by-step calculations.

Introduction to Thermodynamic Properties of Monoatomic Gases

Monoatomic gases, such as helium, neon, argon, and krypton, are characterized by their single-atom molecules. Their thermodynamic behavior is well understood and serves as foundational knowledge in thermodynamics.

Key properties:

- They possess only translational degrees of freedom.
- Their internal energy and enthalpy are primarily due to translational kinetic energy.
- They are ideal gases under standard conditions, simplifying the calculations.

Thermodynamic functions:

- Internal energy (U): Total energy stored within the system.
- Enthalpy (H): Total heat content, defined as $H = U + PV$.
- Temperature dependence: Both U and H depend on temperature, and their changes can be related through specific heat capacities.

Fundamental Concepts and Equations

Internal Energy (U) of a Monoatomic Gas

For an ideal monoatomic gas, the internal energy is given by:

$$ U = \frac{3}{2} nRT $$

Where:

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

The factor $\frac{3}{2}$ corresponds to the three translational degrees of freedom.

Enthalpy (H) of an Ideal Gas

Similarly, enthalpy for an ideal gas relates to its internal energy:

$$H = U + PV$$

Using the ideal gas law $PV = nRT$, it simplifies to:

$$H = U + nRT = \frac{3}{2} nRT + nRT = \frac{5}{2} nRT$$

Hence, for an ideal monoatomic gas:

$$H = \frac{5}{2} nRT$$

This expression indicates that the molar enthalpy depends linearly on temperature.

Change in Enthalpy (ΔH) and Internal Energy (ΔU)

When the temperature changes from T_1 to T_2:

$$
\Delta U = U_2 - U_1 = \frac{3}{2} n R (T_2 - T_1)
$$

$$
\Delta H = H_2 - H_1 = \frac{5}{2} n R (T_2 - T_1)
$$

These relations hold true for ideal gases, where the internal energy and enthalpy are functions solely of temperature.

Converting Temperatures to Kelvin

Since thermodynamic calculations require temperatures in Kelvin:

$$
T(K) = T(°C) + 273.15

\]

Given data:
- Initial temperature \(T_1 = 25°C \Rightarrow 25 + 273.15 = 298.15\,K \)
- Final temperature \(T_2 = 300°C \Rightarrow 300 + 273.15 = 573.15\,K \)

Calculating ΔH for 1 Mole of Monoatomic Gas

Given:
- \(n = 1\, \text{mol} \)
- \(R = 8.314\, \text{J/mol·K} \)
- \(T_1 = 298.15\,K \)
- \(T_2 = 573.15\,K \)

Using the formula:

\[
\boxed{
\Delta H = \frac{5}{2} \times n \times R \times (T_2 - T_1)
}
\]

Plugging in the values:

\[
\Delta H = \frac{5}{2} \times 1 \times 8.314\, \text{J/mol·K} \times (573.15\,K - 298.15\,K)
\]

\[
\Delta H = 2.5 \times 8.314 \times 275
\]

Calculating step-by-step:

1. \(2.5 \times 8.314 = 20.785 \)
2. \(20.785 \times 275 = 5,713.88\, \text{J} \)

Result:

\[
\boxed{
\Delta H \approx 5714\, \text{J}
}
\]

This indicates that the enthalpy of 1 mole of a monoatomic ideal gas increases by approximately 5714 Joules as its temperature rises from 25°C to 300°C.

Deeper Analysis and Additional Considerations

Implications of the Calculation

- The increase in enthalpy corresponds directly to the heat added to the system at constant pressure.
- For monoatomic ideal gases, the relation between temperature and enthalpy change is straightforward and linear.
- This value can be used to analyze processes involving heating, such as in engines, reactors, or calorimetry experiments.

Real Gases and Deviations from Ideal Behavior

While the above calculation assumes an ideal gas, real gases exhibit interactions that can slightly alter the thermodynamic properties:

- Van der Waals corrections introduce parameters to account for finite molecular size and intermolecular forces.
- Heat capacities (Cp and Cv): These can vary with temperature, especially at high temperatures or pressures.
- Non-idealities: For high-pressure conditions, the ideal gas law no longer applies accurately, affecting ΔH calculations.

In such cases, experimental values or equations of state like Van der Waals, Redlich-Kwong, or Peng-Robinson are used to refine calculations.

Heat Capacity Variations with Temperature

Although for monoatomic gases, C_v and C_p are approximately constant over moderate temperature ranges, in real scenarios, these can vary:

- At low temperatures: Quantum effects become significant.
- At high temperatures: Internal degrees of freedom may increase, though for monoatomic gases, no vibrational modes are excited.

The constant volume heat capacity:

$$
C_v = \frac{3}{2} R \approx 12.47\, \text{J/mol·K}
$$

The constant pressure heat capacity:

$$
C_p = \frac{5}{2} R \approx 20.79\, \text{J/mol·K}
$$

Practical Applications and Broader Context

Understanding how to calculate ΔH for monoatomic gases has broad applications:

- Thermal engineering: Design of heating systems, engines, and turbines.
- Calorimetry: Determining heat capacities and energy content.
- Chemical processes: Estimating energy changes during reactions involving gases.
- Astrophysics: Modeling thermodynamic behavior of noble gases in celestial environments.

Summary of Key Takeaways

- For an ideal monoatomic gas, the change in enthalpy when temperature changes from T_1 to T_2 is:

$$
\Delta H = \frac{5}{2} R (T_2 - T_1)
$$

- Converting temperatures to Kelvin is essential for accurate calculations.
- The calculation demonstrates a direct linear relationship between temperature change and enthalpy change for ideal gases.
- Real gases may deviate slightly due to intermolecular forces and variations in heat capacities, but the ideal gas approximation provides a solid baseline.

Conclusion

Calculating the change in enthalpy for one mole of a monoatomic ideal gas over a specified temperature range is a fundamental exercise in thermodynamics that illustrates core principles of energy transfer and molecular behavior. By understanding the relationships between temperature, internal energy, and enthalpy, scientists and engineers can predict and analyze energy flows in various systems, from industrial processes to natural phenomena. While idealized models offer simplicity and clarity, always consider real-world deviations for precise applications. This detailed approach to calculating ΔH provides a robust foundation for mastering thermodynamic calculations involving monoatomic gases.

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