A 14.0 L Container At 323 K Holds A Mixture Of Two Gases With A Total Pressure Of 8.00 Atm. If There

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Understanding the behavior of gases in a mixture is fundamental in chemistry, physics, and engineering. When dealing with gaseous mixtures, key principles such as Dalton's Law of Partial Pressures, the Ideal Gas Law, and mole calculations come into play to determine the properties and composition of the mixture. This article provides a comprehensive overview of how to analyze a gas mixture in a fixed volume and temperature, focusing on calculating individual gas pressures, moles, and the overall behavior of the system.

Introduction to Gas Mixtures and Fundamental Principles

Gases are highly compressible and tend to fill their containers uniformly. When two or more gases occupy the same space, their properties combine in predictable ways governed by physical laws. In this context, understanding how to analyze gas mixtures involves:

- Dalton's Law of Partial Pressures: The total pressure exerted by a mixture of gases equals the sum of the partial pressures of individual gases.
- Ideal Gas Law: $PV = nRT$, relating pressure (P), volume (V), amount of substance in moles (n), temperature (T), and the ideal gas constant (R).

The problem scenario involves a fixed volume and temperature, which simplifies calculations since these parameters remain constant during the process.

Analyzing the Gas Mixture in a Fixed Container

Given data:

- Volume (V) = 14.0 L
- Temperature (T) = 323 K
- Total pressure (P_{total}) = 8.00 atm

Goals:

- Find the number of moles of each gas
- Determine their partial pressures
- Understand the composition of the mixture

Applying the Ideal Gas Law

The Ideal Gas Law offers a straightforward way to relate the measurable quantities:

$$ PV = nRT $$

Where:

- P = pressure in atm
- V = volume in liters
- n = number of moles
- R = ideal gas constant (0.0821 L·atm/mol·K)
- T = temperature in Kelvin

Because the total pressure and volume are known, the total moles of gas in the mixture can be calculated as:

$$ n_{total} = \frac{P_{total} \times V}{R \times T} $$

Plugging in the given values:

$$ n_{total} = \frac{8.00\, \text{atm} \times 14.0\, \text{L}}{0.0821\, \text{L·atm/mol·K} \times 323\, \text{K}} $$

$$ n_{total} \approx \frac{112}{26.53} $$

$$ n_{total} \approx 4.22\, \text{mol} $$

This total molar amount represents the combined moles of the two gases within the container.

Determining Partial Pressures of Each Gas

Dalton's Law states:

$$ P_{total} = P_1 + P_2 $$

Where:

- P_1 = partial pressure of Gas 1

- P_2 = partial pressure of Gas 2

If additional data about the mole fractions or the individual gases is provided, the partial pressures can be calculated directly:

$$P_i = x_i \times P_{total}$$

Where x_i is the mole fraction of gas i:

$$x_i = \frac{n_i}{n_{total}}$$

Example Scenario:
Suppose you know the moles of each gas, say:

- Gas 1: n_1
- Gas 2: n_2

And:

$$n_1 + n_2 = 4.22\, \text{mol}$$

Then:

$$P_1 = \frac{n_1}{n_{total}} \times P_{total}$$

$$P_2 = \frac{n_2}{n_{total}} \times P_{total}$$

Alternatively, if you know the partial pressures or mole fractions, you can derive the individual quantities directly.

Calculating the Moles of Each Gas

Suppose the problem provides the partial pressure of Gas 1 as 3.00 atm, then:

$$n_1 = \frac{P_1 \times V}{RT} = \frac{3.00 \times 14.0}{0.0821 \times 323} \approx 1.55\, \text{mol}$$

Remaining moles for Gas 2:

$$n_2 = n_{total} - n_1 = 4.22 - 1.55 = 2.67\, \text{mol}$$

And its partial pressure:

$$P_2 = \frac{n_2 \times R \times T}{V} = \frac{2.67 \times 0.0821 \times 323}{14.0} \approx 5.00\, \text{atm}$$

Check:

$$ P_1 + P_2 = 3.00 + 5.00 = 8.00\, \text{atm} $$

which matches the total pressure, confirming the calculations.

Implications and Applications of Gas Mixture Analysis

Understanding the composition and behavior of gas mixtures is essential in various fields:

- Industrial Gas Production: For designing reactors and storage tanks.
- Environmental Science: Monitoring atmospheric gases and pollutants.
- Medical Applications: Analyzing inhaled gases and anesthetic mixtures.
- Chemical Engineering: Designing separation processes like gas chromatography.

Practical Considerations:

- Gas behavior may deviate from ideal assumptions at high pressures or low temperatures.
- Real gases require corrections using equations such as Van der Waals' equation.
- Precise measurements of partial pressures require gas sampling and analysis.

Conclusion

Analyzing a mixture of gases in a fixed-volume, fixed-temperature container involves applying the Ideal Gas Law and Dalton's Law of Partial Pressures. By calculating the total moles of gases present, their individual mole fractions, and partial pressures, chemists and engineers can determine the composition and behavior of the mixture accurately.

Understanding these principles is fundamental for designing chemical processes, environmental monitoring, and various industrial applications. With careful calculation and consideration of real-world deviations, the behavior of gaseous mixtures can be predicted and controlled effectively.

Summary of Key Steps:

1. Use the ideal gas law to calculate total moles:

$$ n_{total} = \frac{P_{total} \times V}{R \times T} $$

2. Determine individual moles if mole fractions or partial pressures are known:

$$ n_i = x_i \times n_{total} $$

3. Calculate partial pressures:

$$ P_i = \frac{n_i \times R \times T}{V} $$

4. Verify total pressure:

$$ P_{total} = P_1 + P_2 $$

By following these steps, one can analyze any gas mixture under similar conditions effectively.

Keywords: Gas mixture analysis, Ideal Gas Law, Dalton’s Law, partial pressure, molar calculations, gas behavior, chemical engineering, atmospheric gases, pressure calculations, thermodynamics

Frequently Asked Questions

What is the total pressure exerted by a mixture of gases in a 14.0 L container at 323 K?

The total pressure is 8.00 atm, as given in the problem statement.

How can Dalton's Law of Partial Pressures be applied to this gas mixture?

Dalton's Law states that the total pressure is the sum of the partial pressures of each gas; thus, each gas's partial pressure can be calculated based on its mole fraction and total pressure.

What additional information is needed to determine the individual amounts of each gas in the mixture?

The molar masses of the gases and either the mole ratios or the partial pressures are needed to calculate the individual moles and amounts of each gas.

How can the ideal gas law be used to find the moles of gases in the mixture?

Using PV = nRT, where P is total pressure, V is volume, R is the ideal gas constant, and T is temperature, you can calculate the total moles of gas in the container.

If the gases are ideal, how do their partial pressures relate to

their mole fractions?

The partial pressure of each gas equals its mole fraction multiplied by the total pressure.

What is the significance of temperature (323 K) in this gas mixture problem?

Temperature affects the kinetic energy of the gases and is used in the ideal gas law to determine the number of moles and behavior of the gases.

How can we determine the individual partial pressures of the gases in the mixture?

By knowing the mole fractions or the individual amounts, we multiply each mole fraction by the total pressure (8.00 atm) to find their partial pressures.

What assumptions are made when applying the ideal gas law to this mixture?

Assumptions include that the gases behave ideally, have negligible interactions, and occupy volumes much smaller than the container volume.

How does the volume (14.0 L) influence the calculations for the gas mixture?

The volume, along with pressure and temperature, is used in the ideal gas law to determine the total moles of gas present in the container.

What steps are involved in solving for the individual quantities of each gas in this mixture?

First, calculate the total moles using $PV = nRT$, then determine mole fractions if individual data is available, and finally find partial pressures and individual molar amounts accordingly.

Additional Resources

Understanding Gas Behavior in a Mixture: An In-Depth Analysis of a 14.0 L Container at 323 K with a Total Pressure of 8.00 atm

The behavior of gases in mixtures is a foundational concept in chemistry and physics, offering insights into their properties, interactions, and applications across various scientific and industrial fields. This article explores a specific scenario: a 14.0-liter container held at a temperature of 323 K, containing a mixture of two gases under a total pressure of 8.00 atm. By dissecting this scenario step-by-step, we will delve into the principles governing gas mixtures, the application of the ideal gas law, Dalton's law of partial pressures, and how these concepts converge to provide a comprehensive understanding of the system's behavior.

Fundamentals of Gas Behavior and Mixtures

The Nature of Gases and Their Ideal Behavior

Gases are characterized by their high compressibility, low density, and the ability to expand to fill their containers uniformly. The ideal gas law ($PV = nRT$) models the behavior of gases under many conditions, assuming that gas particles are point masses with no intermolecular forces. While real gases deviate from ideality at high pressures or low temperatures, many laboratory and industrial conditions allow the ideal approximation to be sufficiently accurate.

In a pure gas, the state variables—pressure (P), volume (V), temperature (T), and amount of substance (n)—are directly related. When multiple gases are mixed, their individual behaviors combine in ways described by Dalton's law of partial pressures, which states that each gas exerts its own pressure independently of others.

Gas Mixtures and Their Significance

Gas mixtures are ubiquitous—air being the most common example—comprising primarily nitrogen, oxygen, and trace gases. Understanding these mixtures involves determining:

- The partial pressure of each component.
- The mole fraction of each gas.
- How temperature and volume influence the distribution of gases.

The composition affects properties like reactivity, diffusion rates, and physical behavior, making it essential to analyze the individual contributions within the mixture.

Analyzing the Given System

System Description and Known Parameters

The scenario involves a 14.0 L container at a temperature of 323 K, containing a mixture of two gases with a total pressure of 8.00 atm. The key parameters are:

- Volume (V): 14.0 L
- Temperature (T): 323 K
- Total Pressure (P_{total}): 8.00 atm

The problem statement suggests that further information might be available or needed—such as the composition of the gases, the individual amounts, or the partial pressures. Without explicit details about the individual gases, we can explore the problem using general principles and common assumptions.

Applying the Ideal Gas Law to the Mixture

The ideal gas law relates the total amount of gas (n_total) to the state variables:

$$ PV = nRT $$

Where:

- P = pressure (atm)
- V = volume (L)
- n = moles of gas
- R = ideal gas constant (0.0821 L·atm/(mol·K))
- T = temperature (K)

Calculating the total moles:

$$
n_{total} = \frac{PV}{RT} = \frac{8.00\, \text{atm} \times 14.0\, \text{L}}{0.0821\, \text{L·atm/(mol·K)} \times 323\, \text{K}}
$$

$$
n_{total} = \frac{112}{26.517} \approx 4.22\, \text{mol}
$$

This total mole count provides a basis for subsequent calculations, such as partial pressures if the mole fractions are known.

Partial Pressures and Mole Fractions

The Concept of Partial Pressure

Dalton’s law states that the total pressure of a gas mixture is the sum of the partial pressures of individual gases:

$$ P_{total} = P_1 + P_2 $$

Where:

- P_1 and P_2 are the partial pressures of gases 1 and 2 respectively.

The partial pressure of each gas is proportional to its mole fraction:

$$P_i = x_i \times P_{total}$$

with

$$x_i = \frac{n_i}{n_{total}}$$

Understanding the partial pressures allows us to analyze how each gas contributes to the overall system behavior.

Determining Mole Fractions

Suppose the mixture consists of gases A and B, with mole numbers n_A and n_B. Then:

$$x_A = \frac{n_A}{n_A + n_B}$$

$$x_B = \frac{n_B}{n_A + n_B}$$

The partial pressures:

$$P_A = x_A \times P_{total}$$

$$P_B = x_B \times P_{total}$$

If additional data such as the mole ratio or partial pressures are provided, these calculations can be refined accordingly.

Implications of the System and Potential Further Calculations

Determining the Composition and Behavior of the Gas Mixture

Given the total pressure and temperature, the system's behavior can be further analyzed if the individual gas identities and their mole ratios are known or assumed. For example, if the gases are ideal and non-reactive, their partial pressures and mole fractions remain consistent unless the system undergoes a change.

Common scenarios include:

- Equal molar contributions: Both gases contribute equally to the total pressure.
- Dominance of one gas: One gas's mole fraction is significantly higher, affecting physical and chemical properties.
- Reaction considerations: If gases are reactive, the composition might change over time, influencing the pressure and temperature.

Calculating Partial Pressures in Practice

Suppose the gases are nitrogen (N_2) and oxygen (O_2), common in atmospheric compositions, or other gases like CO_2 and CH_4. If, for example, the mole fraction of gas A is 0.6, then:

$$
P_A = 0.6 \times 8.00\, \text{atm} = 4.80\, \text{atm}
$$
$$
P_B = 0.4 \times 8.00\, \text{atm} = 3.20\, \text{atm}
$$

Using these partial pressures, the individual mole numbers can be calculated via the ideal gas law:

$$
n_i = \frac{P_i V}{RT}
$$

which for each gas provides insights into their respective quantities and potential reactivity.

Practical Applications and Broader Significance

Industrial and Environmental Relevance

Understanding the behavior of gas mixtures under given conditions is vital for numerous applications:

- Chemical manufacturing: Designing reactors and separation processes.
- Environmental science: Modeling atmospheric gases and pollution dispersion.
- Aerospace engineering: Managing pressurized cabins.

- Medical applications: Gas delivery systems.

Accurate calculations of partial pressures and molar quantities impact safety, efficiency, and environmental compliance.

Limitations and Real-World Considerations

While the ideal gas law provides a robust framework, real gases exhibit deviations, especially at high pressures or low temperatures. Corrections such as the Van der Waals equation incorporate molecular size and intermolecular forces, refining predictions.

Furthermore, non-ideal interactions between gases can influence partial pressures, especially if the gases are reactive or form complexes.

Conclusion

Analyzing a 14.0-liter container at 323 K with a total pressure of 8.00 atm containing a mixture of two gases encapsulates fundamental principles of gas behavior, including the ideal gas law, Dalton's law of partial pressures, and mole fraction calculations. These concepts are crucial for understanding not only laboratory systems but also real-world applications across multiple industries.

By systematically applying these principles, scientists and engineers can predict how changes in temperature, pressure, or composition influence the physical state and reactivity of gas mixtures. Such insights facilitate the design of better chemical processes, environmental models, and safety protocols, underscoring the importance of mastering the behavior of gases in mixtures.

Ultimately, this scenario exemplifies the interplay between theoretical frameworks and practical considerations, reinforcing the importance of foundational gas laws in advancing scientific and technological progress.

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