A 1.50 G Sample Of Methanol (ch3oh) Is Placed In An Evacuated 1.00 L Container At 30. Oc. Calculate The

Understanding the Problem: A 1.50 G Sample Of Methanol in an Evacuated Container

A 1.50 G Sample Of Methanol (CH_3OH) Is Placed In An Evacuated 1.00 L Container At 30°C. Calculate The...

While the problem statement appears incomplete, it hints at a typical gas law calculation involving methanol vaporization and pressure determination. This article aims to guide you through the process of calculating the pressure exerted by methanol vapor in the container using the ideal gas law. We will explore concepts such as molar mass, the ideal gas law equation, and the steps involved in solving similar problems.

Introduction to Methanol and Its Properties

What is Methanol?

- Also known as methyl alcohol, CH_3OH.
- A common solvent and fuel additive.
- Has a molar mass of approximately 32.04 g/mol.
- Exhibits both liquid and vapor phases depending on temperature and pressure.

Physical Properties Relevant to the Calculation

- Boiling point at standard atmospheric pressure: approximately 64.7°C.
- At 30°C, methanol exists as a mixture of liquid and vapor phases; vapor pressure is less than atmospheric pressure.

Fundamentals of Gas Laws Relevant to the Calculation

The Ideal Gas Law

- The primary equation used: PV = nRT.
- Variables:
- P = pressure (atm or Pa)
- V = volume (L)
- n = number of moles of gas
- R = ideal gas constant (0.0821 L·atm/mol·K)
- T = temperature in Kelvin (K)

Applying the Ideal Gas Law to Condensed Liquids

- When calculating vapor pressure or the amount of vapor in equilibrium with a liquid, we often assume the vapor behaves ideally.
- The amount of vapor present depends on the temperature and the vapor pressure of the substance.

Step-by-Step Calculation Process

1. Determine the Moles of Methanol in the Sample

- Given mass: 1.50 g
- Molar mass of CH_3OH: approximately 32.04 g/mol

Number of moles:

$$
n = \frac{\text{mass}}{\text{molar mass}} = \frac{1.50\, \text{g}}{32.04\, \text{g/mol}} \approx 0.0468\, \text{mol}
$$

2. Understand the Vapor Pressure of Methanol at 30°C

- Vapor pressure is the pressure exerted by the vapor when in equilibrium with its liquid.
- At 30°C, the vapor pressure of methanol is approximately 100 mm Hg (or 0.1316 atm).

Note: For precise calculations, consult tables or literature data for vapor pressure at different temperatures.

3. Calculate the Partial Pressure of Methanol Vapor

- Since the container is initially evacuated, vaporization occurs until equilibrium vapor pressure is reached.
- The vapor pressure at 30°C is the maximum pressure exerted by the vapor in the container.

Assumption: The vapor phase behaves ideally, and the partial pressure of methanol vapor reaches its vapor pressure at equilibrium.

4. Use the Ideal Gas Law to Confirm or Calculate the Pressure

- Convert temperature to Kelvin:

$$
T = 30°C + 273.15 = 303.15\, K
$$

- Calculate the pressure exerted by the vapor:

$$
P = \frac{nRT}{V}
$$

$$
P = \frac{0.0468\, \text{mol} \times 0.0821\, \frac{L\, atm}{mol\, K} \times 303.15\, K}{1.00\, L}
$$

- Performing the calculation:

$$
P \approx \frac{0.0468 \times 0.0821 \times 303.15}{1}
$$

$$
P \approx 1.16\, \text{atm}
$$

Interpretation: The calculated pressure (~1.16 atm) exceeds the vapor pressure of methanol at 30°C (~0.1316 atm), indicating that not all the methanol vaporizes under these conditions; the vapor pressure limits the vapor pressure at equilibrium.

Implications of Vapor Pressure and Actual Pressure

Vapor Pressure as a Limiting Factor

- Since vapor pressure at 30°C is approximately 0.1316 atm, the maximum pressure exerted by methanol vapor in the container cannot exceed this value at equilibrium.
- The initial calculation suggests that only a fraction of the methanol vaporizes until vapor pressure equilibrium is established.

Calculating the Amount of Vaporized Methanol

- To find the actual amount of vaporized methanol, use the vapor pressure as the equilibrium pressure.
- The number of moles of vapor:

$$
n_{vapor} = \frac{PV}{RT}
$$

$$
n_{vapor} = \frac{0.1316\, \text{atm} \times 1.00\, \text{L}}{0.0821\, \frac{L\, atm}{mol\, K} \times 303.15\, K}
$$

$$
n_{vapor} \approx 0.0053\, \text{mol}
$$

- Corresponding mass of methanol vapor:

$$
m_{vapor} = n_{vapor} \times \text{molar mass} \approx 0.0053\, \text{mol} \times 32.04\, \text{g/mol} \approx 0.17\, \text{g}
$$

Conclusion: Only about 0.17 g of methanol vaporizes at 30°C, and the vapor pressure is approximately 0.1316 atm.

Summary of Key Takeaways

- The initial mass of methanol determines the total potential vapor amount.
- Vapor pressure at a given temperature limits the maximum pressure the vapor can exert.
- Applying the ideal gas law helps estimate the amount of vaporized methanol.
- Actual vapor pressure is often less than the pressure calculated assuming complete vaporization, due to the equilibrium vapor pressure.

Additional Considerations for Accurate Calculations

Non-Ideal Gas Behavior

- At high pressures or low temperatures, gases deviate from ideal behavior.
- Corrections can be applied using the Van der Waals equation.

Realistic Conditions and Experimental Data

- Vapor pressure values vary with temperature; accurate data should be used for precise

calculations.
- Experimental measurements can verify theoretical predictions.

Practical Applications of This Calculation

- Design of chemical reactors involving methanol vaporization.
- Safety assessments related to vapor pressure and pressure buildup.
- Engineering applications involving storage and transport of methanol.

Conclusion

In this article, we've demonstrated how to analyze a problem involving a methanol sample placed in an evacuated container at a specific temperature. By calculating the number of moles, understanding vapor pressure, and applying the ideal gas law, we can estimate the vapor pressure and the amount of methanol vapor present. Such calculations are fundamental in chemical engineering, environmental science, and safety management, ensuring proper handling and understanding of volatile chemicals like methanol.

Further Resources

- "Chemistry: The Central Science" by Brown, LeMay, Bursten.
- NIST Chemistry WebBook for vapor pressure data.
- Online calculators for ideal gas law applications.
- Safety data sheets (SDS) for methanol.

Remember: Always handle chemicals with appropriate safety measures and consult reliable data sources for precise calculations.

Frequently Asked Questions

What is the initial mass of methanol in the sample?

The initial mass of methanol is 1.50 grams, as given in the problem statement.

How do you determine the number of moles of methanol in the sample?

Calculate moles using molar mass: molar mass of CH_3OH is approximately 32.04 g/mol. So, moles = 1.50 g / 32.04 g/mol ≈ 0.0468 mol.

What is the significance of the container being evacuated and at constant temperature?

An evacuated container ensures no external gases influence the system, and constant temperature (30°C) allows the use of the ideal gas law to determine pressure after expansion.

How can the ideal gas law be applied to find the pressure of the methanol vapor?

Use PV = nRT, where P is pressure, V is volume (1.00 L), n is moles of methanol, R is the gas constant (0.0821 L·atm/(mol·K)), and T is temperature in Kelvin (30°C = 303.15 K).

What is the approximate vapor pressure of methanol at 30°C, and why is it relevant?

The vapor pressure of methanol at 30°C is approximately 0.55 atm. This is relevant to determine if the vapor phase is in equilibrium with liquid or if the pressure calculated exceeds the vapor pressure, indicating possible condensation.

How do you interpret the calculated pressure in the context of methanol's vapor-liquid equilibrium?

If the calculated pressure exceeds methanol’s vapor pressure at 30°C, it suggests that not all methanol vaporizes; some may condense. If it is below, the vapor is in equilibrium with the liquid phase.

Additional Resources

A 1.50 G Sample Of Methanol (CH3OH) Is Placed In An Evacuated 1.00 L Container At 30.°C. Calculate The

When examining the behavior of methanol (CH_3OH) in a confined environment, such as a 1.00 L container at 30°C, the fundamental principles of gas laws and thermodynamics come into play. This scenario provides a rich context for exploring various calculations, including the determination of pressure, molar quantities, and the application of ideal gas law principles. Understanding these concepts not only helps in academic settings but also has practical implications in chemical engineering, laboratory safety, and industrial processes. In this comprehensive review, we will dissect the problem step-by-step, explore related concepts, and analyze the broader applications and considerations of handling methanol in sealed environments.

Understanding the Scenario

The scenario involves a small quantity of methanol, a common organic solvent and fuel, introduced into an evacuated (vacuum) container. The key data points provided are:

- Mass of methanol: 1.50 grams
- Volume of container: 1.00 liters
- Temperature: 30°C (which converts to 303.15 K)
- Initial state of the container: evacuated (meaning initially no gas molecules inside)

The primary objective is to calculate the pressure exerted by methanol vapor inside the container at this temperature.

Fundamental Concepts and Principles

Before diving into calculations, it is essential to understand the foundational principles involved:

Ideal Gas Law

The ideal gas law is expressed as:

$$PV = nRT$$

where:
- P = pressure (atm or Pa)
- V = volume (L or m^3)
- n = number of moles
- R = ideal gas constant (0.082057 L·atm/(mol·K))
- T = temperature (K)

This law assumes gases behave ideally, which is a reasonable approximation at moderate pressures and temperatures.

Vapor Pressure and Volatility of Methanol

Methanol is volatile, meaning it readily evaporates at room temperature. Its vapor pressure at 30°C is approximately 97 mm Hg (or about 0.127 atm). This vapor pressure indicates the equilibrium pressure exerted by methanol vapor when the liquid and vapor phases coexist.

Calculating Moles of Methanol

Given the mass and molar mass:

$$ \text{Molar mass of CH}_3\text{OH} = 12.01 + (3 \times 1.008) + 16.00 = 32.042\, \text{g/mol} $$

Number of moles:

$$ n = \frac{\text{mass}}{\text{molar mass}} = \frac{1.50\, \text{g}}{32.042\, \text{g/mol}} \approx 0.0468\, \text{mol} $$

Calculating the Vapor Pressure of Methanol in the Container

Given the vapor pressure at 30°C is approximately 97 mm Hg, it suggests that if methanol is in equilibrium with its vapor phase in the container, the pressure exerted by the vapor would be close to this value. To verify this, we need to determine whether the vapor phase reaches equilibrium under these conditions or if the pressure differs due to the finite amount of methanol.

Step 1: Determine the Moles of Vapor in Equilibrium

Assuming the vapor phase of methanol reaches equilibrium with the liquid phase, the vapor pressure corresponds to the pressure exerted by the vapor at equilibrium.

Using the ideal gas law:

$$ P = \frac{n_{vapor} RT}{V} $$

where:
- n_{vapor} = moles of vapor

If all methanol vaporizes, the maximum number of moles of vapor is limited by the initial amount of methanol.

Since the total moles of methanol are approximately 0.0468 mol, and assuming complete vaporization (which is an overestimate because vapor pressure is partial vaporization), the pressure exerted by the vapor can be calculated.

Step 2: Calculate the Expected Vapor Pressure

Using:

- $n = 0.0468\, \text{mol}$
- $V = 1.00\, \text{L}$
- $T = 303.15\, \text{K}$
- $R = 0.082057\, \text{L·atm/(mol·K)}$

Plugging into the ideal gas law:

$$P = \frac{0.0468 \times 0.082057 \times 303.15}{1.00}$$

Calculating the numerator:

$$0.0468 \times 0.082057 \times 303.15 \approx 1.162\, \text{atm}$$

This indicates that if all methanol vaporized into the 1-liter container, the pressure would be approximately 1.16 atm, which exceeds the vapor pressure at 30°C (~0.127 atm). This discrepancy suggests that under equilibrium conditions, only a small fraction of the methanol would vaporize, and the vapor pressure would be limited to the equilibrium vapor pressure (~0.127 atm).

Thus, the vapor pressure of methanol at 30°C is approximately 97 mm Hg (~0.127 atm).

Implications of Vapor Pressure and Non-Ideal Behavior

The difference between the calculated pressure if all methanol vaporized and the actual vapor pressure emphasizes the importance of vapor-liquid equilibrium. In practical terms:

- Vapor pressure indicates the maximum partial pressure of methanol vapor at equilibrium.
- In the sealed container, if the initial methanol amount is low, the vapor pressure will be close to the equilibrium vapor pressure, and only a small fraction of the methanol will evaporate.
- Complete vaporization is unlikely unless the temperature is higher or the container is larger.

Additional Calculations and Considerations

Beyond simple pressure calculations, several other factors and calculations are relevant:

1. Amount of Methanol Vapor in the Container

Using vapor pressure:

$$ P_{vapor} = 0.127\, \text{atm} $$

Number of moles of vapor:

$$ n_{vapor} = \frac{PV}{RT} = \frac{0.127 \times 1.00}{0.082057 \times 303.15} \approx 0.0052\, \text{mol} $$

This is about 11% of the total methanol (0.0468 mol), indicating that, at equilibrium, only about 11% of the methanol has vaporized.

2. Mass of Methanol in Vapor Phase

Mass of vaporized methanol:

$$ 0.0052\, \text{mol} \times 32.042\, \text{g/mol} \approx 0.167\, \text{g} $$

Remaining liquid methanol:

$$ 1.50\, \text{g} - 0.167\, \text{g} = 1.33\, \text{g} $$

This highlights that most of the methanol remains in the liquid phase under these conditions.

Features, Pros, and Cons of Handling Methanol in Confined Spaces

Understanding the behavior of methanol vapor in sealed containers has practical significance:

Features:
- Methanol is highly volatile with a significant vapor pressure at room temperature.
- It evaporates readily, establishing an equilibrium vapor pressure that depends on temperature.

- Its vapor is flammable and toxic, necessitating careful handling.

Pros:
- Useful as a solvent in laboratory and industrial applications.
- Its vapor pressure allows for easy vapor-phase reactions at controlled pressures.
- Small quantities can be safely contained if proper precautions are taken.

Cons:
- Flammability risk due to methanol vapor accumulation.
- Toxicity poses health hazards upon inhalation or skin contact.
- Vapor pressure varies with temperature, requiring careful control in processes.
- Complete vaporization can lead to pressure buildup, risking container failure if not properly vented.

Additional Applications and Broader Context

The principles demonstrated here extend beyond this specific problem:

- Industrial Storage: Methanol is stored in pressure-rated tanks considering its vapor pressure to prevent accidents.
- Chemical Reactions: Precise control over vapor pressures aids in designing reactions involving methanol vapor.
- Environmental Safety: Understanding vapor-liquid equilibrium assists in designing proper ventilation and containment systems.
- Thermodynamics Education: This problem exemplifies real-world applications of gas laws, vapor pressure, and phase equilibrium.

Conclusion

This detailed exploration of a 1.50 g methanol sample in a sealed 1.00 L container at 30°C illustrates the interplay between physical chemistry principles, practical handling considerations, and safety protocols. The key takeaway is that, under these conditions, the vapor pressure of methanol remains close to its equilibrium vapor pressure (~97 mm Hg), with only a small fraction vaporized. Calculations demonstrate how to estimate vapor quantities, pressures, and the importance of thermodynamic principles in anticipating the behavior of volatile substances. Proper understanding of these

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