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Understanding the properties of gases and how they behave under different conditions is fundamental in chemistry. When given specific data such as mass, volume, pressure, and temperature, we can determine various other characteristics of the gas, such as molar mass, number of moles, or even the gas's identity. In this article, we will analyze the given scenario step-by-step using the ideal gas law, explain the relevant calculations, and explore related concepts to deepen your understanding of gas laws and their applications.

Understanding the Scenario

Before diving into calculations, let's clarify what information we have and what we aim to find:

- Mass of gas: 0.239 grams
- Volume of container: 100 milliliters (mL) or 0.1 liters (L)
- Pressure exerted by gas: 603 millimeters of mercury (mm Hg)
- Temperature: 14°C (which we will convert to Kelvin)

Our goal is typically to find a relevant property such as the molar mass of the gas or the number of moles present under these conditions. The phrase "What Is" implies that we want to find an unknown property, most likely the molar mass, given the data.

Fundamental Concepts and Equations

To analyze the gas, we rely on the ideal gas law:

The Ideal Gas Law

$$PV = nRT$$

Where:
- P = pressure (in atm)
- V = volume (in liters)
- n = number of moles
- R = ideal gas constant ($0.0821 \, \text{L·atm/(mol·K)}$)
- T = temperature (in Kelvin)

This law relates the pressure, volume, temperature, and amount of gas, making it ideal for our calculations.

Converting Units

Since the data provided are in non-standard units, we need to convert them:

1. Pressure from mm Hg to atm
$$ 1\, \text{atm} = 760\, \text{mm Hg} $$
$$ P = \frac{603\, \text{mm Hg}}{760} \approx 0.7934\, \text{atm} $$

2. Temperature from Celsius to Kelvin
$$ T(K) = 14 + 273.15 = 287.15\, \text{K} $$

3. Volume from mL to Liters
$$ V = 100\, \text{mL} = 0.1\, \text{L} $$

Calculating the Number of Moles of Gas

With all units converted, we can now find the number of moles (n).

Step-by-step Calculation

1. Apply the ideal gas law:
$$ n = \frac{PV}{RT} $$

2. Insert the known values:
$$ n = \frac{(0.7934\, \text{atm}) \times (0.1\, \text{L})}{(0.0821\, \text{L·atm/(mol·K)}) \times (287.15\, \text{K})} $$

3. Calculate numerator:
$$ 0.7934 \times 0.1 = 0.07934 $$

4. Calculate denominator:
$$ 0.0821 \times 287.15 \approx 23.565 $$

5. Compute n:
$$ n = \frac{0.07934}{23.565} \approx 0.00337\, \text{mol} $$

Result: The gas sample contains approximately 0.00337 mol.

Determining the Molar Mass of the Gas

Now that we know the number of moles, we can find the molar mass (M):

$$ M = \frac{\text{mass}}{\text{moles}} $$

Given:

- Mass = 0.239 grams
- Moles = 0.00337 mol

Calculating:

$$ M = \frac{0.239\, \text{g}}{0.00337\, \text{mol}} \approx 70.9\, \text{g/mol} $$

Interpretation:
The molar mass of the gas is approximately 70.9 g/mol. This value can be used to identify the gas or compare it to known substances.

Possible Identity of the Gas

Based on the molar mass (~71 g/mol), we can hypothesize potential gases:

- Chlorine dioxide (ClO_2): ~67.45 g/mol
- Nitrogen monoxide (NO): ~30.01 g/mol
- Cyanogen (C_2N_2): ~52 g/mol
- Chlorine (Cl_2): 70.9 g/mol

Given the molar mass close to 71 g/mol, chlorine gas (Cl_2) is a likely candidate. However, further data such as chemical reactivity or spectroscopic analysis would be necessary for precise identification.

Additional Calculations and Considerations

1. Verifying the Ideal Gas Assumption

While the ideal gas law provides a good approximation, real gases may deviate under certain conditions. Since our calculations are based on ideal behavior, factors such as high pressure or low temperature could cause deviations.

2. Impact of Temperature and Pressure

Understanding how temperature and pressure influence gas behavior is critical. For example, increasing temperature at constant volume increases the pressure, whereas increasing pressure at constant temperature can cause deviations from ideality.

3. Molar Mass and Gas Identification

Determining molar mass from experimental data is a common method in qualitative analysis to identify unknown gases.

Practical Applications of Gas Law Calculations

The process we've followed has numerous practical applications:

- Industrial Gas Production: Ensuring proper conditions for manufacturing gases.
- Environmental Monitoring: Analyzing pollutant gases based on their molar masses.
- Laboratory Analysis: Identifying unknown gases in research settings.
- Engineering: Designing systems involving gas storage and transport.

Conclusion

In summary, given a 0.239 g sample of gas in a 100-mL flask exerting a pressure of 603 mm Hg at 14°C, we calculated the molar mass to be approximately 70.9 g/mol. This was achieved by converting all units appropriately, applying the ideal gas law, and performing systematic calculations. The result offers insight into the identity of the gas and demonstrates the power of fundamental gas laws in laboratory and industrial contexts. Understanding these calculations enhances our ability to analyze gases and predict their behavior under various conditions.

Further Reading

For those interested in exploring this topic further, consider reviewing:

- The Ideal Gas Law and its limitations
- Real Gas Behavior and Van der Waals equation
- Techniques for gas identification in analytical chemistry
- Applications of gas laws in environmental science and engineering

By mastering these concepts, you'll be better equipped to handle complex gas-related problems and understand the underlying principles governing gaseous substances.

Frequently Asked Questions

What is the molar mass of the gas in the sample?

The molar mass can be calculated using the ideal gas law: $M = (m \times R \times T) / (P \times V)$. Using $m = 0.239$ g, $V = 0.1$ L, $T = 14°C$ (287 K), $P = 603$ mm Hg (0.793 atm), and $R = 0.0821$ L·atm/(mol·K), the molar mass is approximately 30.1 g/mol.

How do you convert pressure from mm Hg to atmospheres?

To convert mm Hg to atmospheres, divide the pressure in mm Hg by 760. So, 603 mm Hg ÷ 760 ≈ 0.793 atm.

What is the significance of the temperature in Kelvin for gas calculations?

Temperature must be in Kelvin when using the ideal gas law because Kelvin is an absolute scale that starts at zero, ensuring calculations of pressure and volume are accurate and proportional.

Which gas law is most appropriate for calculating the molar mass in this scenario?

The ideal gas law (PV = nRT) is most appropriate, as it relates pressure, volume, temperature, and moles of gas to determine the molar mass.

How does the mass of the gas influence the calculation of molar mass?

The mass of the gas (0.239 g) is used as the numerator in the calculation of molar mass, which is obtained by dividing the total mass by the number of moles (n).

What are common sources of error when calculating molar mass from gas data?

Common errors include incorrect unit conversions, assuming ideal gas behavior when real gases deviate at certain conditions, and measurement inaccuracies in pressure, volume, or temperature.

Additional Resources

A 0.239 Gram Sample Of A Gas In A 100-ml Flask Exerts A Pressure Of 603 Mm Hg At 14 Celsius. What Is — this intriguing problem combines fundamental principles of gas laws, molecular physics, and unit conversions. Whether you're a student tackling chemistry coursework or a professional analyzing gas behaviors in a laboratory, understanding how to interpret and solve such problems is crucial. In this comprehensive guide, we will break down the problem step-by-step, clarify key concepts, and demonstrate how to find the unknowns using the ideal gas law and related principles.

Understanding the Problem Context

The problem provides several pieces of data:

- Mass of the gas: 0.239 grams
- Volume of the container: 100 mL (which is 0.1 liters)
- Pressure exerted by the gas: 603 mm Hg
- Temperature: 14°C

The primary goal is typically to determine an unknown parameter, often the molar mass of the gas, but the problem as presented seems to be incomplete in the snippet. However, based on typical gas law problems, a common question is:

"What is the molar mass of the gas?"

or

"What is the molar amount of the gas?"

or

"How many moles of gas are present?"

In this guide, we will assume the main question is: "Given the data, what is the molar mass of the gas?"

Step 1: Convert All Units to Standard SI Units

Before applying the ideal gas law, it is essential to express all quantities consistently.

Pressure Conversion

- Given: 603 mm Hg
- Conversion factor: 1 atm = 760 mm Hg

Calculate pressure in atmospheres:

$$
P = \frac{603\, \text{mm Hg}}{760\, \text{mm Hg/atm}} \approx 0.7934\, \text{atm}
$$

Temperature Conversion

- Given: 14°C
- Convert to Kelvin:

$$
T(K) = 14 + 273.15 = 287.15\, \text{K}
$$

Volume

- Given: 100 mL
- Convert to liters:

$$
V = 100\, \text{mL} = 0.1\, \text{L}
$$

Mass

- Given: 0.239 grams

Step 2: Apply the Ideal Gas Law

The ideal gas law relates pressure (P), volume (V), number of moles (n), and temperature (T):

$$
PV = nRT
$$

Where:

- P in atm
- V in liters
- n in moles
- $R = 0.0821\, L \cdot atm / mol \cdot K$

Calculate the number of moles (n):

$$
n = \frac{PV}{RT}
$$

Plugging in the known values:

$$
n = \frac{0.7934\, \text{atm} \times 0.1\, \text{L}}{0.0821\, L \cdot atm / mol \cdot K \times 287.15\, K}
$$

Calculating numerator:

$$
0.7934 \times 0.1 = 0.07934
$$

Calculating denominator:

$$
0.0821 \times 287.15 \approx 23.59
$$

Thus,

$$
n \approx \frac{0.07934}{23.59} \approx 0.003365\, \text{mol}
$$

Step 3: Determine the Molar Mass of the Gas

Molar mass (M) is the mass per mole:

$$
M = \frac{\text{mass}}{\text{moles}} = \frac{0.239\, \text{g}}{0.003365\, \text{mol}} \approx 71.0\, \text{g/mol}
$$

Result: The molar mass of the gas is approximately 71 g/mol.

Additional Considerations and Validations

1. Accuracy of Assumptions

- The ideal gas law is an approximation that assumes gases behave ideally at the given conditions.
- For most gases at moderate pressures and temperatures, this approximation is valid.

2. Possible Sources of Error

- Measurement inaccuracies in pressure, temperature, or mass.
- Deviations from ideal behavior, especially at high pressures or low temperatures.

3. Relevance of Molar Mass

Knowing the molar mass helps identify the gas:

- Approximate molar masses:
- Nitrogen (N_2): 28 g/mol
- Oxygen (O_2): 32 g/mol
- Carbon dioxide (CO_2): 44 g/mol
- Propane (C_3H_8): 44 g/mol
- Chlorine (Cl_2): 70.9 g/mol

Given the calculated molar mass (~71 g/mol), the gas could potentially be chlorine (Cl_2).

Summary of Key Steps

Step	Action	Details
1	Convert units	mm Hg to atm, °C to K, mL to L
2	Apply ideal gas law	Calculate number of moles n
3	Compute molar mass	$M = \frac{\text{mass}}{n}$
4	Interpret results	Identify possible gases based on molar mass

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Final Thoughts

This problem exemplifies the practical application of the ideal gas law in laboratory analysis and chemical identification. By carefully converting units, applying the fundamental equation, and interpreting the results, chemists can determine properties like molar mass, which aids in identifying unknown gases or confirming purity.

Understanding gas behavior through such calculations not only enhances theoretical knowledge but also sharpens practical laboratory skills, making you more proficient in chemical analysis and research.

Additional Resources

- Textbooks: "Chemistry: The Central Science" by Brown et al.
- Online Tools: Gas law calculators, unit converters
- Laboratory Tips: Always verify measurement units and calibration of instruments

Feel free to revisit each step, run your own calculations, and explore how variations in pressure, temperature, or mass affect the molar mass and behavior of gases. Mastering these fundamentals opens the door to advanced topics like real gas behavior, reaction kinetics, and thermodynamic processes.

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