What Pressure, In Atm, Would Be Exerted By 0.023 Grams Of Oxygen (O2) If It Occupies 31.6 ML At 91C?

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Understanding the relationship between the amount of gas, its volume, temperature, and pressure is fundamental in chemistry and physics. In this article, we will explore how to determine the pressure exerted by a specific amount of oxygen gas under given conditions. Specifically, we will analyze how 0.023 grams of oxygen (O_2), occupying a volume of 31.6 milliliters at a temperature of 91°C, exerts pressure in atmospheres (atm). Through systematic calculations, we will demonstrate how to approach such problems using the ideal gas law and related concepts.

Understanding the Problem

Before diving into calculations, let’s clarify the problem:

- Mass of oxygen (O_2): 0.023 grams
- Volume occupied: 31.6 mL
- Temperature: 91°C
- Goal: Find the pressure exerted by this oxygen in atm

This problem involves several key concepts:

- Converting mass to moles
- Converting volume to liters
- Converting temperature to Kelvin
- Applying the ideal gas law: PV = nRT

Fundamental Concepts and Constants

To proceed, we need to recall some important constants and conversion factors:

- Molar mass of oxygen (O_2): 32.00 g/mol
- Ideal gas constant (R): 0.082057 atm·L/(mol·K)
- Conversion factors:
- 1 mL = 0.001 L
- Temperature from Celsius to Kelvin: T(K) = T(°C) + 273.15

Step-by-Step Calculation

Let's break down the calculation into manageable steps:

1. Convert mass of oxygen to moles

Number of moles (n) is calculated as:

$n = \text{mass} / \text{molar mass}$

$n = 0.023\ \text{g} / 32.00\ \text{g/mol}$

$n \approx 0.00071875\ \text{mol}$

2. Convert volume to liters

Volume in liters (V):

$V = 31.6\ \text{mL} \times 0.001\ \text{L/mL} = 0.0316\ \text{L}$

3. Convert temperature to Kelvin

$T = 91°\text{C} + 273.15 = 364.15\ \text{K}$

4. Apply the ideal gas law to find pressure

The ideal gas law:

$PV = nRT$

Rearranged for pressure (P):

$P = nRT / V$

Substitute the known values:

$P = (0.00071875\ \text{mol}) \times (0.082057\ \text{atm·L/(mol·K)}) \times (364.15\ \text{K}) / 0.0316\ \text{L}$

Calculate numerator:

$0.00071875 \times 0.082057 \times 364.15 \approx 0.02152$

Calculate pressure:

$P \approx 0.02152 / 0.0316 \approx 0.680$ atm

Therefore, the oxygen exerts approximately 0.68 atm of pressure under the given conditions.

Understanding the Significance of the Calculation

Determining the pressure exerted by a specific amount of gas under certain conditions is crucial in various scientific and industrial applications, including:

- Gas storage and transport
- Chemical reactions involving gases
- Respiratory physiology
- Environmental monitoring

This calculation showcases how the ideal gas law provides a straightforward way to relate measurable quantities like mass, volume, and temperature to the pressure exerted by a gas.

Factors Affecting Gas Pressure

While the ideal gas law offers a simplified model, real gases may deviate due to factors like:

- Intermolecular forces
- Gas particle volume

In most laboratory conditions, these deviations are minimal, and the ideal gas approximation remains valid. However, for high pressures or low temperatures, corrections may be necessary.

Additional Considerations and Practical Applications

Practical Applications of Gas Pressure Calculations

- Designing gas storage tanks: Ensuring tanks can withstand the pressure exerted by stored gases.
- Medical applications: Calculating oxygen pressure in medical devices.
- Environmental science: Estimating atmospheric oxygen levels under different conditions.

Tips for Accurate Calculations

- Always convert all quantities to consistent units.
- Use the correct molar mass for the specific gas.
- Convert temperature to Kelvin to avoid errors.
- Double-check calculations, especially when dealing with small or large values.

Summary of the Calculation Process

Step	Description	Formula / Conversion	Result
1	Convert mass to moles	n = 0.023 g / 32 g/mol	~0.000719 mol
2	Convert volume to liters	31.6 mL × 0.001	0.0316 L
3	Convert Celsius to Kelvin	91 + 273.15	364.15 K
4	Apply ideal gas law	P = nRT / V	~0.68 atm

This structured approach ensures clarity and accuracy in solving gas law problems.

Conclusion

By carefully converting all quantities to appropriate units and applying the ideal gas law, we find that 0.023 grams of oxygen occupying 31.6 mL at 91°C exerts approximately 0.68 atm of pressure. This calculation exemplifies fundamental principles of gas behavior and provides insight into how small quantities of gases can exert measurable pressures under specific conditions. Whether in laboratory settings or industrial applications, mastering these calculations is essential for scientists, engineers, and students alike.

Further Reading and Resources

- "Chemistry: The Central Science" by Brown, LeMay, Bursten
- Online calculators for gas law problems
- Educational videos explaining ideal gas law applications
- Research papers on real gas deviations at high pressures

FAQs

Q1: Can the ideal gas law be used for gases at very high pressures?
At very high pressures, gases may deviate from ideal behavior. In such cases, equations of state like the Van der Waals equation provide more accurate results.

Q2: Why is it important to convert temperature to Kelvin?
Because the ideal gas law is based on absolute temperature scales; Kelvin ensures proportionality and correct calculations.

Q3: How accurate is this calculation for real-world scenarios?
For small gases at moderate conditions, the ideal gas law provides a good approximation. Deviations increase at extreme conditions.

Q4: What is the significance of pressure in physical and chemical processes?
Pressure influences reaction rates, phase changes, and the behavior of gases in various systems.

Q5: How can this calculation be adapted for other gases?
Replace the molar mass and use the specific gas constant if necessary, following the same calculation steps.

By understanding and applying these principles, scientists and students can accurately determine gas pressures under various conditions, enhancing their comprehension of physical chemistry and its applications.

Frequently Asked Questions

How do I calculate the pressure exerted by a given mass of oxygen in a specific volume and temperature?

You can use the ideal gas law: P = (nRT) / V, where n is the number of moles, R is the gas constant, T is temperature in Kelvin, and V is volume in liters.

What is the first step to find the pressure exerted by 0.023 grams of oxygen in this problem?

Convert the mass of oxygen to moles by dividing the mass by the molar mass of O_2 (32 g/mol).

How do I convert the given temperature from Celsius to Kelvin?

Add 273.15 to the Celsius temperature: 91°C + 273.15 = 364.15 K.

What unit should the volume be in to use in the ideal gas law calculation?

Convert the volume from milliliters to liters: 31.6 mL = 0.0316 L.

What is the molar mass of oxygen (O_2), and how is it

used in this calculation?

The molar mass of O2 is approximately 32 g/mol; use it to convert grams to moles by dividing the mass by 32 g/mol.

What is the value of the gas constant R in appropriate units for this calculation?

Use R = 0.0821 L·atm/(mol·K) for calculations involving pressure in atm, volume in liters, temperature in Kelvin, and moles.

What is the final step to find the pressure exerted by the oxygen sample?

Calculate the number of moles, then apply the ideal gas law: P = (nRT) / V, to find the pressure in atm.

Additional Resources

What Pressure, In Atm, Would Be Exerted By 0.023 Grams Of Oxygen (O_2) If It Occupies 31.6 mL At 91°C?

Understanding the pressure exerted by a gas under specific conditions is a fundamental aspect of chemistry, especially when dealing with gases in confined spaces. In this article, we explore the question: what pressure, in atm, would be exerted by 0.023 grams of oxygen (O_2) if it occupies 31.6 mL at 91°C? By walking through the problem step-by-step, applying the ideal gas law, and considering the relevant conversions, we aim to provide a clear, comprehensive guide to solving such a scenario.

Introduction to the Problem

At first glance, this problem involves several key variables:

- Mass of oxygen (O_2): 0.023 grams
- Volume occupied: 31.6 mL
- Temperature: 91°C
- Pressure: To be determined in atm

The challenge lies in connecting these data points through the ideal gas law, which relates pressure, volume, temperature, and moles of gas:

$$ PV = nRT $$

Where:

- P: Pressure (atm)
- V: Volume (L)
- n: Number of moles
- R: Ideal gas constant (0.082057 L·atm/(mol·K))
- T: Temperature (Kelvin)

Step 1: Convert All Quantities to Appropriate Units

Before using the ideal gas law, ensure all units are consistent:

1. Mass to Moles

The molar mass of oxygen (O_2):

- Atomic mass of O: approximately 16.00 g/mol
- O_2: 2 × 16.00 g/mol = 32.00 g/mol

Calculate moles:

$$
n = \frac{\text{mass}}{\text{molar mass}} = \frac{0.023\, \text{g}}{32.00\, \text{g/mol}} \approx 0.00071875\, \text{mol}
$$

2. Volume to Liters

Given volume:

$$
V = 31.6\, \text{mL} = 0.0316\, \text{L}
$$

3. Temperature to Kelvin

Temperature in Celsius:

$$
T_C = 91^\circ C
$$

Convert to Kelvin:

$$
T_K = 91 + 273.15 = 364.15\, \text{K}
$$

Step 2: Apply the Ideal Gas Law to Find Pressure

Rearranged the ideal gas law to solve for pressure:

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

Plug in the known values:

$$
P = \frac{(0.00071875\, \text{mol}) \times 0.082057\, \text{L·atm/(mol·K)} \times 364.15\, \text{K}}{0.0316\, \text{L}}
$$

Calculate numerator:

$$
0.00071875 \times 0.082057 \times 364.15 \approx 0.021567\, \text{L·atm}
$$

Now, divide by the volume:

$$
P = \frac{0.021567}{0.0316} \approx 0.682\, \text{atm}
$$

Step 3: Interpretation of Results

The pressure exerted by 0.023 grams of oxygen occupying 31.6 mL at 91°C is approximately 0.682 atm. This value indicates that under these conditions, the oxygen gas exerts a pressure less than standard atmospheric pressure (1 atm), which makes sense given the small amount of gas and relatively large volume.

Additional Considerations and Real-World Implications

While the ideal gas law provides a straightforward calculation, it's important to consider real-world factors:

1. Deviations from Ideal Gas Behavior

At high pressures or low temperatures, gases deviate from ideal behavior. However, at the conditions specified (a relatively low pressure and moderate temperature), the ideal gas law is sufficiently accurate.

2. Temperature Effects

Increasing temperature would increase the pressure, assuming volume and moles are constant, as per Gay-Lussac's law.

3. Practical Applications

Understanding how to calculate gas pressure under various conditions is essential in fields such as:

- Chemical engineering
- Respiratory physiology (e.g., understanding oxygen pressures in lungs)
- Environmental science (e.g., atmospheric studies)
- Laboratory gas measurements

Summary of the Calculation Process

To determine the pressure exerted by a given amount of gas:

1. Convert the mass of the gas to moles using molar mass.
2. Convert the volume to liters.
3. Convert temperature to Kelvin.
4. Use the ideal gas law \(P = \frac{nRT}{V} \) to find pressure in atm.

Applying these steps meticulously ensures accurate results. In our case, the pressure exerted by 0.023 grams of oxygen in 31.6 mL at 91°C is approximately 0.68 atm.

Final Thoughts

Gases, despite their invisibility, follow predictable physical laws that allow scientists and engineers to calculate their behavior precisely. Whether you're designing a chemical process, studying respiratory function, or just curious about how gases exert pressure under various conditions, mastering the application of the ideal gas law is fundamental. This example demonstrates how a small amount of oxygen can exert a measurable pressure in a confined space, emphasizing the importance of unit conversions and careful calculations in physical chemistry.

Remember: Always verify your conversions and consider the context of your calculations to ensure accuracy and applicability in real-world scenarios.

What Pressure In Atm Would Be Exerted By 0 023 Grams Of Oxygen O2 If It Occupies 31 6 Ml At 91c

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