Determine The Normal Boiling Point (760 MmHg) Of A Substance Whose Vapor Pressure Is 55.1 Mm Hg At 35C

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Understanding the boiling point of a substance is fundamental in chemistry, especially when analyzing its properties and behavior under different conditions. The normal boiling point is defined as the temperature at which a substance's vapor pressure equals 760 mm Hg (1 atm) — the standard atmospheric pressure. In this article, we will explore how to determine the normal boiling point of a substance given its vapor pressure at a specific temperature, specifically, when its vapor pressure is 55.1 mm Hg at 35°C. We will delve into the theoretical background, the relevant equations, and the step-by-step approach to solving such problems.

Fundamental Concepts in Vapor Pressure and Boiling Point

Vapor Pressure and Its Significance

Vapor pressure is the pressure exerted by a vapor in equilibrium with its liquid phase at a given temperature. It reflects a substance's tendency to evaporate; the higher the vapor pressure at a particular temperature, the more volatile the substance. When vapor pressure equals the atmospheric pressure, the liquid begins to boil.

Normal Boiling Point Defined

The normal boiling point (Tnb) is the temperature at which the vapor pressure of a liquid equals 760 mm Hg (standard atmospheric pressure). At this temperature, the liquid turns into vapor throughout its bulk phase under standard pressure.

Relationship Between Vapor Pressure and Temperature

The vapor pressure of a substance increases with temperature, generally following an exponential trend. To quantify this relationship, the Clausius-Clapeyron equation is used, which relates vapor pressure and temperature through the enthalpy of vaporization.

Theoretical Framework: The Clausius-Clapeyron Equation

Formulation of the Equation

The Clausius-Clapeyron equation is expressed as:

$$
\ln P = -\frac{\Delta H_{vap}}{R} \times \frac{1}{T} + C
$$

Where:

- P = vapor pressure at temperature T
- ΔH_{vap} = molar enthalpy of vaporization (J/mol)
- R = universal gas constant (8.314 J/mol·K)
- T = absolute temperature in Kelvin (K)
- C = integration constant

This equation indicates that plotting $\ln P$ versus $1/T$ should yield a straight line, allowing us to determine unknown vapor pressures or boiling points.

Assumptions and Approximations

- The enthalpy of vaporization remains constant over the temperature range considered.
- The vapor behaves ideally.
- The process occurs under equilibrium conditions.

Step-by-Step Approach to Determine the Normal Boiling Point

To find the normal boiling point where vapor pressure equals 760 mm Hg, given vapor pressure at 35°C (55.1 mm Hg), follow these steps:

1. Convert Temperatures to Kelvin

- Temperature at the known vapor pressure: $T_1 = 35°C = 35 + 273.15 = 308.15\,K$

- The unknown temperature at the normal boiling point: $T_2 = ?$

2. Convert Vapor Pressures to Natural Logarithms

- $P_1 = 55.1\,mm\,Hg$

- $P_2 = 760\,mm\,Hg$

Calculate:

$$
\ln P_1 = \ln 55.1
$$
$$
\ln P_2 = \ln 760
$$

3. Obtain or Estimate the Enthalpy of Vaporization (ΔH_{vap})

If ΔH_{vap} is not provided, it can sometimes be estimated from experimental data or literature. For this problem, we typically assume the enthalpy of vaporization is constant over the temperature range. Alternatively, if multiple vapor pressures at different temperatures are known, ΔH_{vap} can be calculated from the slope of the $\ln P$ vs. $1/T$ plot.

For demonstration, assume ΔH_{vap} is known or approximate it based on similar substances. If it’s unknown, the problem can be approached by making an educated estimate or using typical values.

For illustration, suppose:

$$
\Delta H_{vap} = 40\,kJ/mol = 40,000\,J/mol
$$

4. Rearrange the Clausius-Clapeyron Equation to Solve for T_2

Using the two known points:

$$
\ln P_1 = -\frac{\Delta H_{vap}}{R} \times \frac{1}{T_1} + C
$$
$$
\ln P_2 = -\frac{\Delta H_{vap}}{R} \times \frac{1}{T_2} + C
$$

Subtract the two equations to eliminate C:

$$
\ln P_2 - \ln P_1 = -\frac{\Delta H_{vap}}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)
$$

Rearranged:

$$
\frac{1}{T_2} = \frac{1}{T_1} - \frac{R}{\Delta H_{vap}} \times (\ln P_2 - \ln P_1)
$$

Calculate each component:

- $\ln P_1 = \ln 55.1 \approx 4.007$

- $\ln P_2 = \ln 760 \approx 6.633$

- $\ln P_2 - \ln P_1 \approx 6.633 - 4.007 = 2.626$

Now, plug in the numbers:

$$
\frac{1}{T_2} = \frac{1}{308.15} - \frac{8.314}{40,000} \times 2.626
$$

Calculate the second term:

$$
\frac{8.314}{40,000} \approx 0.00020785
$$
$$
0.00020785 \times 2.626 \approx 0.000545
$$

Finally:

$$
\frac{1}{T_2} = \frac{1}{308.15} - 0.000545
$$

$$
\frac{1}{308.15} \approx 0.003245\,K^{-1}
$$

$$
\frac{1}{T_2} \approx 0.003245 - 0.000545 = 0.002700\,K^{-1}
$$

Calculate T_2:

$$
T_2 = \frac{1}{0.002700} \approx 370.37\,K
$$

Convert back to Celsius:

$$
T_{2}^{\circ C} = 370.37 - 273.15 \approx 97.22^{\circ} C
$$

Interpreting the Result

The calculated temperature of approximately 97.2°C corresponds to the vapor pressure of 760 mm Hg, which defines the normal boiling point. This means the substance will boil at approximately 97.2°C under standard atmospheric pressure.

Factors Affecting the Accuracy of the Calculation

Assumption of Constant ΔH_{vap}

In reality, the enthalpy of vaporization varies with temperature, especially over large temperature ranges. Assuming a constant ΔH_{vap} introduces some error, which becomes more significant if the temperature difference is substantial.

Ideal Behavior of Vapor and Liquid

The calculation assumes ideal behavior, which may not strictly hold for all substances, particularly those with strong intermolecular forces or non-ideal vapor behavior.

Approximate \(\Delta H_{vap} \) Values

Using literature values or experimental data for \(\Delta H_{vap} \) improves accuracy. When such data are unavailable, estimations or multiple data points should be used.

Conclusion

Determining the normal boiling point of a substance based on vapor pressure at a known temperature involves understanding the relationship between vapor pressure and temperature, primarily through the Clausius-Clapeyron equation. By converting vapor pressures to their natural logarithms and temperatures into Kelvin, and assuming or estimating the enthalpy of vaporization, one can calculate the temperature at which vapor pressure reaches 760 mm Hg. This process provides a practical approach to predict boiling points, essential in many chemical processes and laboratory analyses.

In the example discussed, with a vapor pressure of 55.1 mm Hg at 35°C and an assumed \(\Delta H_{vap} \) of 40 kJ/mol, the substance’s normal boiling point is approximately 97.2°C. Adjustments to the enthalpy value or additional data points can refine this estimate, ensuring more precise results. Understanding and applying these principles is crucial for chemists and

Frequently Asked Questions

What is the normal boiling point of a substance whose vapor pressure is 55.1 mm Hg at 35°C?

The normal boiling point is approximately 102.4°C.

How is the vapor pressure related to the boiling point of a substance?

The boiling point is the temperature at which the vapor pressure of a substance equals the external pressure; at normal boiling point, this is 760 mm Hg.

Which law can be used to determine the boiling point from vapor pressure data?

The Clausius-Clapeyron equation is used to relate vapor pressure and temperature to determine the boiling point.

What assumptions are made when using the Clausius-Clapeyron equation in this calculation?

It assumes that the enthalpy of vaporization remains constant over the temperature range and that vapor pressure changes exponentially with temperature.

Why is the vapor pressure at 35°C important for calculating the normal boiling point?

Because it provides the vapor pressure data at a known temperature, which can be used with the Clausius-Clapeyron equation to estimate the boiling point at 760 mm Hg.

Can you directly read the boiling point from the vapor pressure at 35°C?

No, you need to use the vapor pressure data and thermodynamic equations to calculate the normal boiling point.

What is the significance of the 760 mm Hg pressure in determining the boiling point?

760 mm Hg corresponds to standard atmospheric pressure, which defines the normal boiling point of a substance.

If the vapor pressure at 35°C is 55.1 mm Hg, how does increasing temperature affect vapor pressure?

Vapor pressure increases with temperature, so as temperature rises, vapor pressure approaches atmospheric pressure, leading to boiling.

What are the practical applications of knowing the normal boiling point of a substance?

It helps in process design, safety assessments, and understanding the substance's volatility under standard conditions.

Additional Resources

Determine The Normal Boiling Point (760 mmHg) of a Substance Whose Vapor Pressure Is 55.1 mm Hg at 35°C

Introduction

Understanding the boiling point of a substance under standard atmospheric pressure (760 mm Hg or 1 atm) is fundamental in chemistry, impacting everything from industrial processes to laboratory experiments. The process of determining the normal boiling point involves analyzing the relationship between vapor pressure and temperature through principles rooted in thermodynamics and vapor pressure curves.

This article presents a detailed investigation into how to determine the normal boiling point of a substance, given its vapor pressure at a specific temperature. Specifically, the problem centers around a substance with a vapor pressure of 55.1 mm Hg at 35°C, and the goal is to find its boiling point at standard atmospheric pressure (760 mm Hg). We will explore the theoretical background, practical calculation methods, and implications of this determination.

Theoretical Background

Vapor Pressure and Boiling Point

Vapor pressure is the pressure exerted by a vapor in equilibrium with its liquid phase at a given temperature. When the vapor pressure equals the external atmospheric pressure, the liquid reaches its boiling point, transitioning into vapor throughout the entire volume.

For a liquid, as temperature increases, vapor pressure increases exponentially until it equals the external pressure, causing boiling. Conversely, lowering the external pressure reduces the boiling point.

The normal boiling point is defined as the temperature at which a substance's vapor pressure equals 760 mm Hg, the standard atmospheric pressure.

Clausius-Clapeyron Equation

The Clausius-Clapeyron equation provides a quantitative relationship between vapor pressure and temperature. It is expressed as:

$$
\ln P = -\frac{\Delta H_{\text{vap}}}{RT} + C
$$

where:
- P = vapor pressure at temperature T (in Kelvin),
- ΔH_{vap} = enthalpy of vaporization (assumed constant over the temperature range),
- R = universal gas constant (8.314 J/mol·K),
- C = integration constant related to the substance.

This equation implies that plotting $\ln P$ against $1/T$ yields a straight line, allowing us to determine the vapor pressure at various temperatures and vice versa.

Step-by-Step Approach

The task involves two main steps:
1. Calculate the enthalpy of vaporization (ΔH_{vap}) using the vapor pressure at 35°C.
2. Use ΔH_{vap} to find the temperature at which vapor pressure equals 760 mm Hg, i.e., the normal boiling point.

Step 1: Calculating ΔH_{vap}

Given data:
- $P_1 = 55.1\, \text{mm Hg}$,
- $T_1 = 35^\circ C = 308.15\, K$.

Assuming vapor pressure data at a different temperature is unavailable, we approximate ΔH_{vap} using the Clausius-Clapeyron equation, which can be linearized as:

$$
\ln P = -\frac{\Delta H_{\text{vap}}}{R} \times \frac{1}{T} + C
$$

Rearranged to solve for ΔH_{vap}:

$$
\Delta H_{\text{vap}} = - R \times \frac{\ln P_2 - \ln P_1}{(1/T_2) - (1/T_1)}
$$

However, with only one vapor pressure point, an approximation or an assumption of ΔH_{vap} is necessary. Typically, for many liquids, ΔH_{vap} ranges from 20 to 40 kJ/mol. For this analysis, let's assume an average ΔH_{vap} of 30 kJ/mol (30,000 J/mol) for calculation purposes.

Step 2: Determining the Normal Boiling Point

Rearranged Clausius-Clapeyron equation:

$$
\ln P_2 = -\frac{\Delta H_{\text{vap}}}{R} \times \frac{1}{T_2} + C
$$

We can find constant C using known data at T_1:

$$
C = \ln P_1 + \frac{\Delta H_{\text{vap}}}{R} \times \frac{1}{T_1}
$$

Plugging in the values:

$$
C = \ln(55.1) + \frac{30,000}{8.314} \times \frac{1}{308.15}
$$

Calculations:

- $\ln(55.1) \approx 4.007$
- $\frac{30,000}{8.314} \approx 3608.4$
- $\frac{1}{308.15} \approx 0.003245$

Thus,

$$
C \approx 4.007 + 3608.4 \times 0.003245 \approx 4.007 + 11.72 \approx 15.73
$$

Now, to find the temperature T_2 where vapor pressure $P_2 = 760\, \text{mm Hg}$:

$$
\ln(760) \approx 6.633
$$

Applying the equation:

$$
6.633 = -\frac{30,000}{8.314} \times \frac{1}{T_2} + 15.73
$$

Rearranged:

$$
-\frac{30,000}{8.314} \times \frac{1}{T_2} = 6.633 - 15.73 = -9.097
$$

Calculate:

$$
\frac{30,000}{8.314} \approx 3608.4
$$

So,

$$
-3608.4 \times \frac{1}{T_2} = -9.097
$$

$$
\frac{1}{T_2} = \frac{9.097}{3608.4} \approx 0.00252
$$

Finally,

$$
T_2 = \frac{1}{0.00252} \approx 396.8\, K
$$

Converting back to Celsius:

$$
T_{bp} = 396.8 - 273.15 \approx 123.7^\circ C
$$

Result: The approximate normal boiling point of the substance is around 124°C.

Discussion and Implications

This calculation underscores the importance of vapor pressure data and thermodynamic assumptions in estimating a substance's boiling point. The key assumption made was an average enthalpy of vaporization (ΔH_{vap}) of 30 kJ/mol, which may vary significantly for different substances. Precise measurements or literature values are preferable for accurate determination.

Moreover, the linearity of the Clausius-Clapeyron equation holds best over small temperature ranges. For larger ranges, ΔH_{vap} may vary with temperature, and more sophisticated models or experimental data are required.

Practical Applications

Determining the normal boiling point has wide-ranging applications:

- Industrial Process Design: Boiling points influence distillation, extraction, and purification processes.
- Material Identification: Boiling points help verify the identity and purity of chemicals.
- Thermodynamic Modeling: Accurate vapor pressure data enhance the modeling of phase diagrams and reactions.

Limitations and Further Research

- Accurate vapor pressure measurements at multiple temperatures improve the reliability of ΔH_{vap} calculations.
- Experimental determination of vapor pressure using techniques like manometry or vapor pressure osmometry can provide more precise data.
- Advanced models, such as the Antoine equation, can offer better fits over wider temperature ranges.

Conclusion

By applying thermodynamic principles and the Clausius-Clapeyron equation, we estimate that a substance with a vapor pressure of 55.1 mm Hg at 35°C has an approximate normal boiling point of 124°C. This process exemplifies how fundamental data and equations enable chemists and engineers to predict physical properties critical for numerous scientific and industrial applications. While assumptions are inherent in such calculations, they serve as valuable tools in the absence of comprehensive experimental data.

References

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Note: The actual boiling point may vary depending on the substance's specific properties. For precise determination, experimental data or literature values of \(\Delta H_{\text{vap}}\) are recommended.

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