Calculate The Final Temperature Of 32 ML Of Ethanol Initially At 11C Upon Absorption Of 562J Of Heat.

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Understanding how substances respond to heat absorption is fundamental in thermodynamics, a branch of physics that deals with heat, work, and energy transfer. In particular, calculating the change in temperature of a liquid when a specific amount of heat is added is a common problem faced by chemists, engineers, and students alike. This type of calculation helps in designing chemical processes, understanding environmental impacts, and even in everyday applications like cooking or refrigeration.

In this article, we will explore how to determine the final temperature of 32 milliliters (mL) of ethanol initially at 11°C after it absorbs 562 joules (J) of heat. Ethanol, a widely used alcohol in various industries, has unique thermodynamic properties that influence how its temperature changes upon heat absorption. By understanding these properties and applying fundamental principles of heat transfer, we can accurately calculate the final temperature.

Understanding the Basics of Heat Transfer and Specific Heat Capacity

Before diving into the calculation, it's essential to review some key concepts:

What is Heat?

Heat is a form of energy transfer between systems or objects due to a temperature difference. When heat is added to a substance, it can cause an increase in temperature, a change in phase, or a chemical reaction, depending on the conditions and the amount of heat.

Specific Heat Capacity (c)

The specific heat capacity of a substance is the amount of heat required to raise the temperature of one gram of the substance by one degree Celsius (or Kelvin). It is expressed in units of J/(g·°C).

Mathematically:

\[
Q = mc\Delta T
\]

Where:

- \(Q\) = heat added (J)
- \(m\) = mass of the substance (g)
- \(c\) = specific heat capacity (J/g·°C)
- \(\Delta T\) = change in temperature (°C)

This equation is fundamental to calculating temperature changes resulting from heat transfer.

Properties of Ethanol Relevant to the Calculation

To accurately perform our calculation, we need the specific heat capacity of ethanol and its density to convert volume to mass.

Specific Heat Capacity of Ethanol

At room temperature (~20°C), the specific heat capacity of ethanol is approximately:

- $c_{ethanol} = 2.44 \, \text{J/g·°C}$

This value can vary slightly with temperature, but for most practical calculations, 2.44 J/g·°C is sufficiently accurate.

Density of Ethanol

Density determines how much mass is contained in a given volume:

- $\rho_{ethanol} \approx 0.789 \, \text{g/mL}$

Thus, 1 mL of ethanol weighs approximately 0.789 grams.

Step-by-Step Calculation of the Final Temperature

The problem involves several steps:

1. Convert volume to mass
2. Apply the heat transfer formula to find the temperature change
3. Calculate the final temperature

Let's analyze each step in detail.

1. Convert Volume to Mass

Given:

- Volume of ethanol, $V = 32\, \text{mL}$
- Density of ethanol, $\rho = 0.789\, \text{g/mL}$

Calculating mass:

$$
m = V \times \rho = 32\, \text{mL} \times 0.789\, \text{g/mL} \approx 25.248\, \text{g}
$$

Round for simplicity:

$$
m \approx 25.25\, \text{g}
$$

2. Calculate the Temperature Change (ΔT)

Using the heat transfer formula:

$$
Q = mc \Delta T
$$

Rearranged to find ΔT:

$$
\Delta T = \frac{Q}{mc}
$$

Substitute the known values:

- $Q = 562\, \text{J}$
- $m = 25.25\, \text{g}$
- $c = 2.44\, \text{J/g·°C}$

Calculation:

$$
\Delta T = \frac{562\, \text{J}}{25.25\, \text{g} \times 2.44\, \text{J/g·°C}} \approx \frac{562}{61.61} \approx 9.12\, °C
$$

This indicates that the ethanol's temperature will increase by approximately 9.12°C upon absorbing 562 J of heat.

3. Determine the Final Temperature

Initial temperature:

$$
T_{initial} = 11°C
$$

Final temperature:

$$
T_{final} = T_{initial} + \Delta T = 11°C + 9.12°C \approx 20.12°C
$$

Summary of the Calculation

Parameter	Value
Volume of ethanol	32 mL
Density of ethanol	0.789 g/mL
Mass of ethanol	25.25 g

Initial temperature	11°C
Heat absorbed	562 J
Specific heat capacity	2.44 J/g·°C
Temperature increase	9.12°C
Final temperature	approximately 20.12°C

Implications and Practical Applications

Knowing how to calculate temperature changes due to heat absorption allows scientists and engineers to design better thermal systems, optimize chemical reactions, and predict the behavior of liquids in various conditions. For instance:

- In chemical processing, understanding temperature changes helps prevent overheating or underheating of reactants.
- In environmental science, predicting how liquids respond to heat impacts ecological models.
- In everyday scenarios like cooking or refrigeration, such calculations ensure safety and efficiency.

Furthermore, this method can be extended to other liquids and solids by substituting appropriate specific heat capacities and densities.

Limitations and Assumptions

While the calculation provides a good estimate, it’s essential to recognize some assumptions:

- Constant Specific Heat Capacity: We assume the specific heat capacity of ethanol remains constant over the temperature range. In reality, it varies slightly with temperature.
- No Heat Loss: The calculation assumes all the heat energy goes into raising the temperature of ethanol, with no heat lost to surroundings. In practical situations, heat loss can occur.
- Pure Substance: The calculations assume pure ethanol without impurities, which can affect heat capacity.

Considering these factors ensures more accurate modeling in real-world applications.

Conclusion

Calculating the final temperature of a liquid after heat absorption is a straightforward process grounded in fundamental thermodynamics principles. For 32 mL of ethanol initially at 11°C absorbing 562 J of heat, the temperature increases by approximately 9.12°C, resulting in a final temperature of around 20.12°C. This calculation hinges on knowing the specific heat capacity and density of ethanol, converting volume to mass, and applying the heat transfer formula.

Mastering such calculations is crucial for professionals and students working in fields ranging from chemical engineering to environmental science. Understanding the energy transfer processes empowers better control and prediction of thermal behavior in various systems.

Keywords: ethanol, heat transfer, specific heat capacity, temperature calculation, thermodynamics, thermal properties, energy transfer, chemical engineering, temperature change, heat absorption

Frequently Asked Questions

How do you calculate the final temperature of ethanol after absorbing heat?

You use the formula $Q = mc\Delta T$, where Q is the heat absorbed, m is the mass, c is the specific heat capacity, and ΔT is the change in temperature. Rearranged, $\Delta T = Q / (m\ c)$.

What is the specific heat capacity of ethanol used in calculations?

The specific heat capacity of ethanol is approximately 2.44 J/g°C.

How do you convert the volume of ethanol to mass for this calculation?

Since the density of ethanol is about 0.789 g/mL, multiply the volume (32 mL) by this density to get the mass: 32 mL × 0.789 g/mL ≈ 25.25 g.

What is the step-by-step process to find the final temperature of ethanol in this problem?

First, convert volume to mass, then calculate $\Delta T = Q / (m\ c)$. Finally, add ΔT to the initial temperature (11°C) to find the final temperature.

What is the final temperature of ethanol after absorbing 562 J of heat?

Using the values: m ≈ 25.25 g, c = 2.44 J/g°C, Q = 562 J, $\Delta T = 562 / (25.25 \times 2.44) \approx 9.2$°C. Therefore, the final temperature is approximately 11°C + 9.2°C = 20.2°C.

Why is it important to use the correct specific heat capacity when calculating temperature change?

Because the specific heat capacity varies for different substances, using the correct value ensures accurate calculation of the temperature change when heat is absorbed or released.

Additional Resources

Thermodynamics: Calculating the Final Temperature of Ethanol After Heat Absorption

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When exploring the fascinating world of thermodynamics, one of the fundamental concepts is understanding how a substance's temperature changes when it absorbs or releases heat. This principle not only underpins scientific research but also plays a crucial role in industries ranging from pharmaceuticals to energy production. Today, we delve into a practical, hands-on problem: Calculating the final temperature of 32 mL of ethanol initially at 11°C after it absorbs 562 joules of heat. This task exemplifies the core ideas of heat transfer, specific heat capacity, and temperature change, providing both an educational and real-world application.

Understanding the Problem Context

Before jumping into the calculations, it's essential to grasp the problem's parameters and what physical principles are involved. The scenario involves:

- Substance: Ethanol (ethyl alcohol)
- Initial Volume: 32 mL

- Initial Temperature: 11°C
- Heat Absorbed: 562 J

The goal is to determine the final temperature of the ethanol after it absorbs this heat.

Fundamental Concepts in Play

To approach this problem systematically, we need to understand the key thermodynamic concepts involved:

2.1 Specific Heat Capacity (c)

Specific heat capacity is a property of a substance that indicates how much heat energy is needed to raise the temperature of a unit mass of the substance by 1°C (or 1 Kelvin). For ethanol, this value is well-documented in scientific literature.

Value of specific heat capacity of ethanol:

- $c_{ethanol} \approx 2.44\, \text{J/g°C}$

This means that to raise 1 gram of ethanol by 1°C, 2.44 joules of energy are needed.

2.2 Mass and Volume Relationship

Since the problem provides volume in milliliters, converting volume to mass is necessary because specific heat capacity is expressed per gram. This requires knowing the density of ethanol.

Density of ethanol:

- At 20°C, approximately $\rho_{ethanol} = 0.789\, \text{g/mL}$

Assuming the temperature change is small and density doesn't vary significantly in this range, this value can be used for approximation.

2.3 Heat Transfer Calculation

The fundamental relationship linking heat transfer and temperature change is:

$$
Q = mc\Delta T
$$

Where:

- Q is the heat energy absorbed or released (in joules)
- m is the mass of the substance (in grams)
- c is the specific heat capacity (J/g°C)
- ΔT is the change in temperature (final - initial)

Step-by-Step Calculation Process

Let's now proceed with the calculation, carefully considering each step.

3.1 Convert Volume to Mass

Given:

- Volume, $V = 32\, \text{mL}$
- Density of ethanol, $\rho = 0.789\, \text{g/mL}$

Calculate mass:

$$
m = V \times \rho = 32\, \text{mL} \times 0.789\, \text{g/mL} = 25.248\, \text{g}
$$

3.2 Calculate Temperature Change (ΔT)

Rearranged from the heat transfer equation:

$$
\Delta T = \frac{Q}{mc}
$$

Substituting the known values:

$$
\Delta T = \frac{562\, \text{J}}{25.248\, \text{g} \times 2.44\, \text{J/g°C}}
$$

Calculating denominator:

$$
25.248\, \text{g} \times 2.44\, \text{J/g°C} \approx 61.613\, \text{J/°C}
$$

Thus,

$$
\Delta T \approx \frac{562}{61.613} \approx 9.12\, °C
$$

3.3 Determine Final Temperature

Since the heat is absorbed, the temperature increases:

$$
T_{final} = T_{initial} + \Delta T
$$

Given:

- $T_{initial} = 11^\circ C$

Therefore:

$$
T_{final} = 11^\circ C + 9.12^\circ C \approx 20.12^\circ C
$$

Interpreting the Results and Practical Considerations

The calculation indicates that after absorbing 562 J of heat, 32 mL of ethanol initially at 11°C will reach approximately 20.12°C. This seemingly straightforward calculation encapsulates a lot of underlying science and assumptions.

4.1 Assumptions in the Calculation

- Constant Specific Heat Capacity: The calculation assumes the specific heat capacity remains constant over the temperature range. While generally valid for small temperature changes, for larger shifts, slight variations could occur.
- Negligible Heat Loss: The model assumes all the heat energy is absorbed solely by the ethanol without any losses to the surroundings, which is idealized.
- Constant Density: Density is assumed constant; in reality, it varies slightly with temperature changes, but this variation is minor over the temperature range (~11°C to ~20°C).

4.2 Practical Implications

Understanding such calculations is essential in various practical applications:

- Chemical Reactions: Precise temperature control during reactions involving ethanol.
- Cooling and Heating Systems: Designing systems that account for heat transfer involving ethanol as a solvent or fuel.
- Laboratory Procedures: Accurate thermal measurements and adjustments to ensure experimental accuracy.

Additional Considerations for Accurate Calculations

While the above method provides a solid estimate, there are additional factors that could refine the calculation:

- Temperature-dependent Specific Heat: For more precise work, use specific heat capacities at the specific temperature range.
- Heat Loss to Environment: Incorporate heat loss factors if conducting the experiment in an open

environment.
- Container Effects: Consider the container's thermal properties, which might absorb or conduct heat away from the ethanol.

Summary

Calculating the final temperature of ethanol after heat absorption involves fundamental thermodynamic principles, primarily the relationship $Q = mc\Delta T$. By converting volume to mass using the known density, applying the specific heat capacity, and performing the calculation, we found that:

> The ethanol's temperature increases by approximately 9.12°C, reaching a final temperature of about 20.12°C.

This exercise underscores the importance of understanding material properties, unit conversions, and the assumptions underlying thermodynamic calculations—a vital skill for scientists and engineers working with thermal systems.

Final Notes

Mastering such calculations enhances your ability to predict thermal behavior in real-world applications, whether in laboratory settings, industrial processes, or product development. Remember, always consider the assumptions made and the context of your calculations to ensure accuracy and applicability. As thermodynamics continues to be a cornerstone of physical sciences, honing these skills is invaluable for advancing research and technological innovation.

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