If This Problem Is Solved Assuming (wrongly) That All The Ice Melts, A Final Temperature Of T = ?16.5C

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Understanding the behavior of ice and water during phase changes is a fundamental aspect of thermodynamics, with numerous practical applications in climate science, engineering, and everyday life. When analyzing problems involving the melting of ice and subsequent temperature changes, it's crucial to account for all relevant energy exchanges, including latent heats and specific heats of the substances involved. However, sometimes in educational settings or simplified models, assumptions are made that do not fully capture the physics of the situation. One such assumption is to consider that all the ice melts completely and then ignore the subsequent temperature adjustments of the resulting water.

This approach can lead to incorrect results, such as predicting a final temperature of T = −16.5°C, which may seem plausible at first glance but is fundamentally flawed if the problem's initial conditions and energy balances are not properly analyzed. In this article, we will explore the reasoning behind such assumptions, why they are incorrect, and how to correctly approach problems involving ice melting and temperature changes.

Understanding the Problem Context

In many thermodynamics problems, you are asked to determine the final temperature after a certain amount of ice is mixed with warmer water or exposed to a heat source. The typical scenario involves:

- A known mass of ice at a known initial temperature (usually 0°C).
- A known mass of water at a higher initial temperature.
- The goal: find the final equilibrium temperature after heat exchange.

A common mistake is to assume that all the ice melts instantly and then to analyze the temperature change of the resulting water without considering the energy required for melting or the heat lost or gained during the phase change.

The Flawed Assumption: All Ice Melts and Final Temperature is −16.5°C

Suppose someone assumes that:

- All the ice melts completely.
- The temperature of the resulting mixture settles at −16.5°C.

This assumption is problematic because it neglects the following:

- The energy required to melt the ice (latent heat of fusion).
- The fact that water cannot exist below 0°C unless it is under pressure or supercooled.
- The conservation of energy principles, which dictate that the heat lost by warmer water must equal the heat gained by the ice and the resulting water.

Assuming the final temperature is −16.5°C suggests that the water cools below 0°C, which is physically impossible unless the water is under pressure or in a supercooled state. Therefore, this conclusion indicates an oversimplification or misapplication of thermodynamic principles.

Why Is This Assumption Incorrect?

1. Ignoring the Latent Heat of Fusion

The latent heat of fusion (L_f) for ice is approximately 334 kJ/kg. When ice melts, it absorbs this energy without changing temperature. If you ignore this, you underestimate the energy needed to convert ice at 0°C into liquid water at 0°C.

Implication: Assuming all ice melts instantly at the same temperature without considering the latent heat leads to an inaccurate calculation of the energy exchange.

2. Neglecting the Temperature Constraints of Liquid Water

Water cannot exist stably below 0°C in an open system, as it freezes unless kept under pressure or supercooled. Therefore, final temperatures below 0°C (such as −16.5°C) are physically impossible for freely existing water.

Implication: The assumption ignores the fundamental thermodynamic fact that the equilibrium temperature after melting and mixing cannot be below 0°C unless under special conditions.

3. Violating Conservation of Energy

Any heat lost by the warmer water must go into melting the ice and raising the temperature of the resulting water. If the initial conditions do not provide enough heat to melt all the ice and warm the mixture to −16.5°C, then that final temperature cannot be achieved.

Implication: The assumption violates the energy conservation principle, leading to physically impossible results.

Correct Approach to Solving the Problem

To accurately determine the final temperature after melting ice and mixing with warmer water, follow these steps:

Step 1: Define Known Quantities

- Mass of ice: \(m_{ice} \)
- Initial temperature of ice: \(T_{ice, initial} = 0°C \) (assuming ice at melting point)
- Mass of water: \(m_{water} \)

- Initial temperature of water: $T_{water, initial}$ (greater than 0°C)

Step 2: Write the Energy Balance Equation

The total heat lost by the warm water equals the heat gained by the ice and the water formed from melted ice:

$$
Q_{lost} = Q_{gained}
$$

Expressed mathematically:

$$
m_{water} c_{water} (T_{water, initial} - T_{final}) = m_{ice} L_f + m_{ice} c_{water} (T_{final} - 0°C)
$$

where:

- c_{water} = specific heat capacity of water (~4.18 kJ/kg·K)
- L_f = latent heat of fusion (~334 kJ/kg)
- T_{final} = final temperature of the mixture (unknown)

Note: If T_{final} is below 0°C, this indicates supercooling or other special conditions, which are generally not considered in standard problems.

Step 3: Solve for T_{final}

Rearranged:

$$
m_{water} c_{water} (T_{water, initial} - T_{final}) = m_{ice} L_f + m_{ice} c_{water} T_{final}
$$

$$
m_{water} c_{water} T_{water, initial} - m_{water} c_{water} T_{final} = m_{ice} L_f + m_{ice} c_{water} T_{final}
$$

Group terms:

$$
m_{water} c_{water} T_{water, initial} - m_{ice} L_f = T_{final} (m_{water} c_{water} + m_{ice} c_{water})
$$

Solve for T_{final}:

$$
T_{final} = \frac{m_{water} c_{water} T_{water, initial} - m_{ice} L_f}{(m_{water} + m_{ice}) c_{water}}
$$

This formula provides the physically meaningful equilibrium temperature considering all energy exchanges.

Implications of the Correct Solution

Applying the correct approach often results in a final temperature at or above 0°C, consistent with physical laws. If the calculation yields a temperature below 0°C, it indicates that not all the ice melts or that the initial assumptions need revision.

In scenarios where the initial heat energy is insufficient to melt all the ice and raise the temperature above 0°C, the system will reach an equilibrium with some ice remaining unmelted, and the temperature will be exactly 0°C.

Common Misconceptions and How to Avoid Them

- Assuming all ice melts instantly without considering latent heat: Always include the latent heat in calculations.
- Ignoring the temperature limits of water: Remember, under normal pressure, water cannot exist below 0°C in liquid form.
- Overlooking energy conservation: Ensure that the heat lost by warmer water equals the heat gained by ice and meltwater.
- Calculating final temperature below 0°C without pressure considerations: Recognize that such conditions are physically unrealistic in open systems.

Conclusion: Why the Assumption of −16.5°C is Flawed

The assumption that, after solving the problem, the final temperature is −16.5°C by considering that all the ice melts is flawed because it neglects the fundamental thermodynamic principles:

- Energy required for melting (latent heat)
- Physical constraints on water temperature
- Conservation of energy

Proper analysis must incorporate the latent heat of fusion, the initial conditions, and the physical limits of water temperature. When correctly performed, the calculations typically show that the final temperature will be at or above 0°C unless under special pressure conditions. Any result indicating a temperature like −16.5°C suggests an oversimplification or misapplication of thermodynamics.

Key Takeaways:

- Always include latent heat in phase change calculations.
- Recognize physical limits of water temperature.
- Use energy balance equations to ensure physically valid results.
- Be cautious with assumptions that oversimplify complex processes.

By understanding these principles, students and professionals can accurately analyze thermodynamic problems involving ice melting and avoid errors like assuming a final temperature of −16.5°C without proper justification.

Meta Description:
Learn why assuming all ice melts and the final temperature is −16.5°C is incorrect in thermodynamics problems.

Discover the proper approach to calculating final temperatures after ice melting, including energy balances and physical constraints.

Frequently Asked Questions

What assumptions are made when calculating the final temperature assuming all the ice melts?

The calculation assumes that all the ice melts completely, that the heat exchange occurs without losses, and that the final temperature T = -16.5°C is reached when the system reaches thermal equilibrium.

Why might assuming all the ice melts lead to an inaccurate final temperature estimate?

Because in real scenarios, some ice may not melt completely or heat transfer may be incomplete, leading to a different final temperature; assuming complete melting simplifies the calculation but can overlook these factors.

How does the specific heat capacity of water influence the final temperature after ice melting?

The specific heat capacity determines how much energy is needed to change water's temperature, affecting the final equilibrium temperature once the ice has melted and mixed with other substances.

What role does the latent heat of fusion play in calculating the final temperature when all ice melts?

Latent heat of fusion is the energy required to turn ice into liquid water without changing temperature; it is critical in energy balance calculations to determine how much heat is absorbed during melting.

Is the temperature T = -16.5°C realistic for a system where all ice melts, and why might this value be considered 'wrongly' assumed?

Typically, melting ice cools the system to around 0°C, so assuming a final temperature of -16.5°C is unrealistic unless additional cooling factors are involved; hence, it's considered a wrongly assumed value.

How can inaccuracies in the initial temperatures of the components affect the final temperature calculation?

Incorrect initial temperature assumptions can lead to errors in energy balance calculations, resulting in an inaccurate estimate of the final temperature, especially if the initial temperatures are significantly different.

What are the practical implications of assuming the final temperature T = -16.5°C in real-world scenarios?

Assuming such a low final temperature could mislead engineers or scientists about the thermal behavior of systems involving ice and water, potentially leading to design flaws or incorrect interpretations of thermal processes.

How should one properly approach calculating the final temperature when all ice melts in a mixture?

One should perform an energy balance considering the heat required to melt the ice (latent heat) and the heat exchanged to reach thermal equilibrium, accurately accounting for initial temperatures and heat capacities.

Additional Resources

Ice Melting and Final Temperature: Analyzing the Assumption That All the Ice Melts to Reach T = -16.5°C

Introduction

The study of thermal processes involving phase changes, particularly the melting of ice and subsequent temperature adjustments, is fundamental in thermodynamics. In many educational and practical contexts, simplified assumptions are made to facilitate calculations—one such assumption being that all ice melts completely before the system reaches thermal equilibrium. While convenient, this assumption can sometimes lead to inaccuracies if the actual physical scenario differs.

In this review, we explore the implications of assuming that all the ice melts, specifically analyzing a hypothetical final temperature T = -16.5°C. We will examine whether this assumption holds true under typical conditions, the physical principles involved, and the consequences of such an approximation. The goal is to understand the depth of this assumption, its validity, and the detailed thermodynamic reasoning behind the actual behavior of the system.

Context and Typical Scenario

Suppose we have a system consisting of:

- An initial quantity of ice at a temperature below 0°C.
- A given amount of warm water or other heat source.
- An insulated environment to prevent heat exchange with surroundings.

The common problem involves determining:

- The final temperature after heat exchange.
- Whether all the ice melts or remains partially solid.
- The phase and temperature state of the system at equilibrium.

In many textbook problems, the assumption that all ice melts simplifies calculations because it allows one to focus solely on the heat required for melting (latent heat) plus the heating of the resulting water to the final temperature.

The Assumption: All Ice Melts to Reach T = -16.5°C

The specific assumption under scrutiny is that, even after heat exchange, the system reaches a final temperature of T = -16.5°C, with the entire ice mass melted and equilibrated at this temperature.

Key points of this assumption:

- The ice has fully melted, turning into liquid water at the final temperature.
- No residual ice remains in solid form.

- The entire system (water and melted ice) is at thermal equilibrium at T = -16.5°C.

Critical Analysis of the Assumption

1. Physical Plausibility of the Final Temperature

The first question to ask is: Is T = -16.5°C a physically consistent final temperature?

- If all ice melts, the system's temperature cannot be below 0°C unless it is a mixture of water and ice in equilibrium, which is common at the melting point.
- At temperatures below 0°C, the system would be in a mixed phase (ice + water), with the ice in partial or complete solid form.
- For all ice to be melted and the system to settle at -16.5°C, the water must be cooled below 0°C, which would only be possible if the water's temperature drops below its melting point, implying the presence of ice still in the system or the need for supercooling.

Conclusion: It is physically inconsistent for a system with all ice melted to have a final temperature below 0°C unless the system contains supercooled water or other anomalous conditions, which are not typical in standard thermodynamic problems.

2. Thermodynamic Constraints

According to the principles of thermodynamics:

- Heat transfer occurs from warmer parts of the system to cooler parts.
- Phase change (melting) requires input of latent heat, L_f, at 0°C.
- Cooling of water below 0°C requires removal of heat, but liquid water cannot exist stably below 0°C unless supercooled.

Implication: To reach a final temperature of -16.5°C, either:

- The water must be supercooled, which is a metastable state and not typical in equilibrium calculations.
- Or the initial assumptions are incorrect, and some ice remains unmelted.

Therefore, the assumption that all ice melts and the system stabilizes at T = -16.5°C is physically inconsistent under normal conditions.

Thermodynamic Calculations and the Correct Approach

1. Energy Balance Equations

In typical problems involving melting ice and temperature change, the energy conservation equation is:

$$
Q_{in} = Q_{out}
$$

where:

- Q_{in}: Heat supplied (e.g., from warm water).
- Q_{out}: Heat lost (e.g., to cold surroundings or initial ice).

When considering phase change:

\[

$$Q_{melting} = m_{ice} \times L_f$$

and for temperature change:

$$Q_{sensible} = m \times c \times \Delta T$$

where:

- m: mass of water or ice.
- c: specific heat capacity.
- L_f: latent heat of fusion (~334 kJ/kg for ice).

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2. Typical Solution Steps

- Calculate heat required to raise ice from initial temperature $T_{ice, initial}$ to 0°C.
- Add the latent heat to melt all ice.
- Use remaining heat to warm the melted ice (now water) from 0°C to the final temperature T_{final}.

Mathematically:

$$Q_{total} = m_{ice} c_{ice} (0 - T_{ice, initial}) + m_{ice} L_f + m_{water} c_{water} (T_{final} - 0)$$

The system reaches equilibrium when the heat supplied equals Q_{total}.

Implications of the Assumption on Final Temperature

Assuming that all ice melts and the system reaches T = -16.5°C:

- Contradicts physical laws because the liquid water cannot exist stably at -16.5°C without freezing or supercooling.
- Suggests an underlying misconception—possibly the assumption is made for the purpose of an idealized calculation, ignoring phase stability.

Realistic Final Temperatures

In reality, the final temperature:

- Is above 0°C if all ice melts and the system reaches equilibrium.
- Can be below 0°C if some ice remains unmelted, resulting in a phase equilibrium at 0°C, with residual ice.

The Role of Supercooling and Supercooling Limitations

Supercooling is a metastable state where water remains liquid below its freezing point, but:

- It is not a stable equilibrium.
- The system will tend to freeze spontaneously if nucleation sites are available.

In thermodynamic equilibrium, the final temperature of a melted ice-water system is at or above 0°C unless supercooling conditions are artificially maintained.

Thus, assuming a final temperature of -16.5°C for a system where all the ice is melted indicates an unphysical scenario unless supercooling is explicitly incorporated, which complicates the analysis significantly.

Practical Considerations and Common Mistakes

- Assuming all ice melts without considering the initial temperature of the ice.
- Ignoring the phase change heat: neglecting the latent heat L_f can lead to incorrect conclusions about the temperature.
- Misinterpreting the final temperature: assuming the final equilibrium temperature can be below 0°C without accounting for the phase state.

Common errors include:

- Taking the final temperature as given without checking phase stability.
- Using simplified formulas without considering the phase diagram constraints.
- Assuming supercooled states are stable equilibria in standard conditions.

Summary and Conclusions

- The assumption that all ice melts and the system reaches a final temperature of T = -16.5°C is physically inconsistent under normal thermodynamic conditions.
- The phase stability of water implies that liquid water cannot persist below 0°C unless supercooling or other special conditions are present.
- In typical systems, if all ice melts, the final temperature is at or above 0°C.
- To accurately determine the final temperature, one must perform an energy balance considering:

- Initial temperatures.
- Phase change heat (latent heat).
- Specific heats.
- Whether residual ice remains.

- The incorrect assumption simplifies calculations but misrepresents the physical reality. It may be useful as an approximation in some contexts, but only when explicitly justified.

Final Remarks

When approaching problems involving phase change and temperature equilibria, it is critical to:

- Use phase diagrams and thermodynamic principles.
- Recognize the limits of assumptions.
- Cross-verify whether the assumptions align with physical laws.

Understanding the nuances of melting, supercooling, and phase stability ensures accurate and meaningful thermodynamic analyses, preventing oversimplifications that lead to unphysical conclusions such as a final temperature of -16.5°C after complete melting of ice.

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Melts A Final Temperature Of T 16 5c

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