A 21 G Block Of Ice Is Cooled To 77 C. It Is Added To 593 G Of Water In An 92 G Copper Calorimeter At

A 21 G Block Of Ice Is Cooled To 77 C. It Is Added To 593 G Of Water In An 92 G Copper Calorimeter At a specific initial temperature, initiating a fascinating exploration of heat transfer, specific heat capacities, and temperature equilibrium. This scenario exemplifies fundamental principles of thermodynamics, including heat exchange between different substances, the role of calorimeters in experimental physics, and the calculations involved in determining final temperatures and heat transfer quantities. Understanding this process not only enhances comprehension of thermal physics but also provides insights into real-world applications such as climate science, refrigeration systems, and material science.

Understanding the Scenario: Key Components and Initial Conditions

Before delving into calculations and concepts, let's clarify the initial conditions and components involved in this thermal process:

- Ice Block: Mass = 21 g; initial temperature = 77°C (note: this is unusual as ice typically melts at 0°C, so this likely refers to a different phase or a conceptual problem)
- Water: Mass = 593 g; initial temperature = unspecified (assumed to be at a certain temperature, often room temperature, say 25°C for calculations)
- Copper Calorimeter: Mass = 92 g; initial temperature = same as water or specified
- Final Equilibrium Temperature: To be determined based on heat exchange

Note: The initial temperature of the water and the copper calorimeter needs to be specified or assumed for precise calculations. For this scenario, we will assume the water and calorimeter are initially at 25°C, a common room temperature.

Fundamental Principles of Heat Transfer in Thermodynamics

This problem revolves around key thermodynamic principles:

1. Conservation of Energy

- The total heat lost by the warmer substances equals the total heat gained by the colder substances, assuming no heat loss to the surroundings.

2. Specific Heat Capacity

- The amount of heat required to raise the temperature of 1 gram of a substance by 1°C.
- Different substances have different specific heat capacities:
- Water: 4.18 J/g°C
- Copper: 0.385 J/g°C
- Ice (solid water): approximately 2.09 J/g°C

3. Phase Changes and Latent Heat

- When a substance undergoes a phase change, heat is either absorbed or released without a change in temperature.
- For ice melting: latent heat of fusion = 334 J/g

Step-by-Step Calculation of Final Equilibrium Temperature

To analyze the heat exchange, we will consider:

- Heating or cooling of the ice to its melting point if necessary
- Melting of ice (if initial temperature is below 0°C)
- Heating of melted ice to the final temperature
- Heating of water and calorimeter to the final temperature

Given the initial condition that the ice is at 77°C, which is above its melting point, this suggests the ice is actually in a different phase or in a scenario where the problem involves cooling the ice from 77°C to a lower temperature, or perhaps there's a typo. Assuming the scenario intends for the ice to be at 0°C initially (which is typical in such problems), the calculations proceed accordingly.

For this analysis, we'll assume:

- Ice initially at 0°C (melting point)
- Water initially at 25°C
- Calorimeter initially at 25°C

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1. Heat Lost or Gained by Ice

- If the ice is initially at 0°C, it can absorb heat to melt and then warm up further.
- The heat required to melt the ice:

$Q_{melt} = m_{ice} \times L_f = 21\,g \times 334\,J/g = 702\,J$

- The heat required to warm the melted ice from 0°C to the final temperature T_f:

$Q_{warm\,ice} = m_{ice} \times c_{water} \times (T_f - 0)$

Note: If the initial temperature of the ice is higher than 0°C, the process will involve cooling the ice first, which alters the calculations.

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2. Heat Lost or Gained by Water and Calorimeter

- The water and calorimeter will either lose or gain heat depending on the final temperature.
- Heat lost or gained by water:

$Q_{water} = m_{water} \times c_{water} \times (T_{initial\,water} - T_f)$

- Heat lost or gained by calorimeter:

$Q_{cal} = m_{cal} \times c_{cal} \times (T_{initial\,cal} - T_f)$

Calculating the Final Temperature (T_f)

By applying conservation of energy:

$$
Q_{lost} + Q_{gained} = 0
$$

or

$$
Q_{water} + Q_{cal} + Q_{ice} = 0
$$

Depending on the initial conditions, the calculation involves:

- Melting the ice (if necessary)
- Heating the melted ice to T_f

- Cooling or heating water and calorimeter to T_f

Implications of Heat Transfer in Practical Applications

Understanding this thermal exchange process is crucial in numerous fields:

1. Climate Science

- Ice melting and heat absorption are central to climate models predicting sea level rise and global temperature changes.

2. Refrigeration and Climate Control

- Devices rely on heat transfer principles similar to this scenario to cool environments efficiently.

3. Material Science

- Knowledge of specific heat capacities and phase changes informs the development of new materials and thermal management systems.

Advanced Topics: Real-World Considerations and Experimental Design

In practical experiments, several factors influence heat transfer calculations:

- Heat losses to surroundings
- Non-ideal thermal contact between substances
- Variations in specific heat capacities with temperature
- Latent heat considerations during phase changes

Designing experiments requires meticulous calibration, insulation, and consideration of all heat flows to obtain accurate data.

Summary of Key Points

- Conservation of energy governs heat transfer processes during mixing and heating.
- Specific heats and latent heats are fundamental to calculating temperature changes.
- The phase of the ice (melting or sublimating) plays a crucial role in energy exchange.
- Calorimeters are essential tools for isolating and measuring thermal effects in experiments.
- Real-world applications extend from climate models to engineering systems.

Conclusion

The scenario of a 21 g block of ice cooled to 77°C and added to water in a copper calorimeter provides a rich context for exploring fundamental thermodynamic principles. Through understanding heat capacities, phase changes, and conservation of energy, we can accurately predict the final temperature of the system and quantify heat transfer. Mastery of these concepts is vital for scientists and engineers working in fields that involve thermal management, climate science, and materials research. As we continue to refine our understanding of heat transfer mechanisms, the insights gained from such experiments contribute to technological advances and a deeper grasp of the natural world.

For further reading on thermodynamics, heat transfer, and calorimetry, consult physics textbooks or reputable online educational resources.

Frequently Asked Questions

What is the initial temperature of the ice block in the problem?

The initial temperature of the ice block is 77°C, as stated in the problem.

How much water is used in the calorimeter setup?

The water used is 593 grams.

What is the mass of the copper calorimeter in this experiment?

The copper calorimeter has a mass of 92 grams.

What is the significance of the ice block's mass being 21 g in this context?

The mass of the ice block (21 g) is used to calculate heat exchange during melting and warming processes.

Why is the specific heat capacity of copper important in this problem?

The specific heat capacity of copper is crucial because it determines how much heat the calorimeter absorbs or releases during temperature changes.

What is the likely final temperature of the system after the ice melts and reaches thermal equilibrium?

The final temperature can be calculated by setting heat lost by water and copper equal to heat gained by melting and warming the ice.

Which principle underlies the calculations in this problem?

The principle of conservation of energy, specifically heat transfer, underlies the calculations.

How can you determine if all the ice melts in this setup?

By comparing the heat available from the water and copper to the heat required to melt and warm the ice; if sufficient, all ice melts; otherwise, only part melts.

Additional Resources

A 21 G Block Of Ice Is Cooled To 77°C. It Is Added To 593 G Of Water In An 92 G Copper Calorimeter At

This intriguing scenario presents a classic thermodynamics problem that involves understanding heat transfer, specific heat capacities, phase changes, and the principles governing thermal equilibrium. Analyzing such a problem offers valuable insights into the fundamental concepts of heat exchange, the role of materials in thermal processes, and the practical applications of calorimetry.

Introduction to the Problem

At its core, the problem involves adding a block of ice, initially cooled to -77°C (assuming the negative sign indicates sub-zero temperature), into a water sample contained within a copper calorimeter. The goal is to understand how heat transfers among the ice, water, and the calorimeter, leading to a thermal equilibrium state. This scenario is emblematic of calorimetry experiments, which are pivotal in physics and chemistry for measuring specific heat capacities, latent heats, and for understanding energy conservation during phase changes.

Understanding the Components

1. The Ice Block

- Mass: 21 grams
- Initial Temperature: -77°C (assuming the negative sign indicates below freezing point)
- State: Solid (ice)
- Properties:
- Specific heat capacity of ice: approximately 2.1 J/g°C
- Latent heat of fusion for ice: about 334 J/g

2. The Water

- Mass: 593 grams
- Initial Temperature: Not specified explicitly, but commonly assumed to be at or near room temperature (e.g., 20°C) for typical problems unless otherwise stated.
- Properties:
- Specific heat capacity of water: 4.18 J/g°C

3. The Copper Calorimeter

- Mass: 92 grams
- Initial Temperature: Same as water, typically assumed to be 20°C unless specified.
- Properties:
- Specific heat capacity of copper: approximately 0.385 J/g°C

Key Assumptions and Considerations

- The system is isolated, with no heat exchange with the environment.
- The ice is initially at -77°C, which is significantly below 0°C; in reality, ice at such a low temperature would need to be cooled from its melting point down to -77°C, which involves cooling via refrigeration.
- The ice will undergo phase change (melting) at 0°C if it reaches that temperature.

- The final temperature of the system, once equilibrium is reached, will be somewhere between the initial temperatures of the components.
- The heat capacities and latent heats are constant over the temperature ranges considered.

Analyzing the Thermodynamics

Step 1: Heating the Ice from -77°C to 0°C

The initial step involves raising the temperature of the ice from -77°C to 0°C.

Heat needed to warm ice:
$$ Q_1 = m_{ice} \times c_{ice} \times \Delta T $$
$$ Q_1 = 21\,g \times 2.1\,J/g°C \times (0 - (-77))°C $$
$$ Q_1 = 21 \times 2.1 \times 77 $$
$$ Q_1 = 21 \times 161.7 $$
$$ Q_1 \approx 3395.7\,J $$

This is the energy required to bring the ice up to 0°C, assuming it is cooled from its initial temperature.

Step 2: Melting the Ice at 0°C

Once at 0°C, the ice begins to melt into water.

Heat for melting:
$$ Q_2 = m_{ice} \times L_f $$
$$ Q_2 = 21\,g \times 334\,J/g $$
$$ Q_2 = 7014\,J $$

The energy required for complete melting is substantial.

Step 3: Heating the Water from 0°C to Final Equilibrium Temperature

Post-melting, the water formed will need to be heated from 0°C to the final temperature T_f.

$$ Q_3 = m_{water} \times c_{water} \times (T_f - 0) $$
$$ Q_3 = 21\,g \times 4.18\,J/g°C \times T_f $$

$$ Q_3 \approx 87.78 \times T_f $$

Step 4: Heating the Water and the Copper Calorimeter to Final Temperature

The existing water and the calorimeter will cool down or warm up depending on the heat exchange.

Heat lost or gained by water:
$$ Q_{water} = m_{water} \times c_{water} \times (T_{initial\,water} - T_f) $$
$$ Q_{water} = 593\,g \times 4.18\,J/g°C \times (T_{initial\,water} - T_f) $$

Heat lost or gained by the copper calorimeter:
$$ Q_{cal} = m_{cal} \times c_{Cu} \times (T_{initial\,cal} - T_f) $$
$$ Q_{cal} = 92\,g \times 0.385\,J/g°C \times (T_{initial\,cal} - T_f) $$

Assuming initial temperatures are 20°C, the heat exchange balances as the system reaches equilibrium.

Setting Up the Energy Balance

Since the system is isolated, the total heat lost by the warmer components equals the total heat gained by the colder components:

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

In practice, the ice gains heat to reach 0°C, melts, and possibly warms further, while the water and calorimeter may lose heat as they cool down to the final temperature T_f.

The energy balance is complex because the ice at -77°C must be heated through multiple stages, and the final temperature depends on the initial conditions and the heat exchanges.

Estimating the Final Temperature

Given the large temperature difference, the dominant heat exchange will involve the ice reaching 0°C and melting, absorbing a significant amount of energy. The water and calorimeter will cool down accordingly.

A simplified approach involves assuming the ice melts completely and the system reaches thermal equilibrium at a temperature T_f somewhere between the initial temperatures.

Approximate process:
- The ice absorbs about 3395.7 J to reach 0°C.
- It then absorbs 7014 J to melt.
- The total energy absorbed by the ice to become water at 0°C is about 10409.7 J.

The thermal energy available from the water and calorimeter is:

$$Q_{water\,and\,cal} = (m_{water} \times c_{water} + m_{cal} \times c_{Cu}) \times (T_{initial} - T_f)$$

$$= (593 \times 4.18 + 92 \times 0.385) \times (20 - T_f)$$
$$\approx (2480.54 + 35.37) \times (20 - T_f)$$
$$\approx 2515.91 \times (20 - T_f)$$

Setting the heat gained by ice equal to the heat lost by water and calorimeter:

$$10409.7\,J + \text{(heat to warm melted ice)} = 2515.91 \times (20 - T_f)$$

If we assume the melted ice warms from 0°C to the final temperature T_f:

$$Q_{warm\,water} = 21\,g \times 4.18\,J/g°C \times (T_f - 0) \approx 87.78 \times T_f$$

Adding this to the melting energy:

$$10409.7 + 87.78 \times T_f = 2515.91 \times (20 - T_f)$$

Expanding:

$$10409.7 + 87.78 T_f = 50318.2 - 2515.91 T_f$$

Bringing all terms to one side:

$$87.78 T_f + 2515.91 T_f = 50318.2 - 10409.7$$
$$2603.69 T_f = 39908.5$$
$$T_f \approx \frac{39908.5}{2603.69} \approx 15.33°C$$

This suggests the final temperature is approximately 15.3°C, indicating the system equilibrates with the water cooling slightly and the ice melting completely.

Features and Insights

Pros of this analysis:
- Demonstrates the application of conservation of energy.

- Clarifies the multi-stage heat transfer process.
- Provides a practical example of calorimetry principles.
- Highlights the importance of phase change and latent heat in thermal calculations.

Cons and limitations:
- Assumes no heat loss to surroundings, which is idealized.
- Uses constant specific heats, neglecting temperature dependence.
- Assumes initial water and calorimeter are at the same temperature.
- Ignores the possibility of incomplete melting

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