

A 42.65 G Sample Of A Substance Is Initially At 25.1 C. After Absorbing 476 Cal Of Heat, The Temperature

A 42.65 G Sample Of A Substance Is Initially At 25.1 C. After Absorbing 476 Cal Of Heat, The Temperature is a classic problem in thermodynamics and calorimetry. It involves understanding how heat transfer affects the temperature of a substance, which is fundamental in fields ranging from chemistry and physics to engineering and environmental science. This scenario provides an excellent opportunity to explore concepts such as specific heat capacity, heat transfer, and the calculations necessary to determine temperature changes in materials.

In this detailed article, we will examine the steps involved in solving this problem, discuss the underlying principles, and explore related concepts to deepen your understanding of heat transfer processes. Whether you're a student preparing for exams or a professional looking to refresh your knowledge, this comprehensive guide aims to clarify the essential ideas and calculations involved.

Understanding the Problem: The Basics of Heat Transfer and Specific Heat

Before diving into calculations, it's important to understand the key concepts at play:

What is Heat Transfer?

Heat transfer refers to the process of thermal energy moving from one object or substance to another due to a temperature difference. When a substance absorbs heat, its internal energy increases, often resulting in a rise in temperature.

What is Specific Heat Capacity?

Specific heat capacity (often denoted as c) is a physical property of a substance that indicates how much heat energy is needed to raise the temperature of one gram of the substance by one degree Celsius (or Kelvin). The SI unit for specific heat capacity is $J/(g \cdot ^\circ C)$, but in calorimetry, calories per gram per degree Celsius ($cal/g \cdot ^\circ C$) are also common.

Key formula:

$$Q = mc\Delta T$$

Where:

- Q = heat absorbed or released (in calories or joules)
- m = mass of the substance (in grams)

- c = specific heat capacity (in $\text{cal/g}\cdot^\circ\text{C}$ or $\text{J/g}\cdot\text{K}$)
- ΔT = change in temperature (in $^\circ\text{C}$ or K)

Step-by-Step Solution Approach

The problem provides:

- Mass of the substance: $m = 42.65\text{ g}$
- Initial temperature: $T_i = 25.1^\circ\text{C}$
- Heat absorbed: $Q = 476\text{ cal}$

Our goal:

- Find the final temperature T_f after heat absorption.

The key unknown in this scenario is the specific heat capacity c of the substance. To proceed, we typically need this value or information to determine it. If the problem provides the specific heat capacity, we can directly calculate the temperature increase. If not, we may need to infer or assume it based on the substance's identity.

Assumption: Known Specific Heat Capacity

Suppose the substance is water, which has a specific heat capacity of approximately $1.00\text{ cal/g}\cdot^\circ\text{C}$. This is a common assumption in calorimetric problems unless specified otherwise.

Calculation:

1. Calculate the temperature change (ΔT):

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

2. Determine the final temperature:

$$T_f = T_i + \Delta T$$

Calculating the Temperature Change

Using the values:

- $Q = 476\text{ cal}$
- $m = 42.65\text{ g}$

- $(c = 1.00, \text{cal/g}\cdot^{\circ}\text{C})$ (assuming water)

Compute:

$$\Delta T = \frac{476 \text{ cal}}{42.65 \text{ g} \times 1.00 \text{ cal/g}\cdot^{\circ}\text{C}} = \frac{476}{42.65} \approx 11.16, ^{\circ}\text{C}$$

Final Temperature:

$$T_f = 25.1, ^{\circ}\text{C} + 11.16, ^{\circ}\text{C} \approx 36.26, ^{\circ}\text{C}$$

Thus, after absorbing 476 calories, the temperature of the substance increases from 25.1°C to approximately 36.26°C .

Understanding the Significance of Specific Heat Capacity

If the specific heat capacity of the substance differs from that of water, the temperature change will vary accordingly. For example:

Substance	Approximate Specific Heat Capacity (cal/g $\cdot^{\circ}\text{C}$)	Note
Water	1.00	Standard reference
Aluminum	0.22	Metals have lower specific heats
Iron	0.11	Very good conductor, low specific heat

Implication: The lower the specific heat capacity, the larger the temperature change for the same amount of heat absorbed.

Example: If the substance is aluminum with $(c \approx 0.22, \text{cal/g}\cdot^{\circ}\text{C})$

Calculate:

$$\Delta T = \frac{476}{42.65 \times 0.22} \approx \frac{476}{9.383} \approx 50.75, ^{\circ}\text{C}$$

Final temperature:

$$T_f = 25.1 + 50.75 \approx 75.85, ^{\circ}\text{C}$$

This illustrates how the material's properties significantly influence temperature change.

Additional Factors and Real-World Considerations

While the idealized calculation provides a good estimate, real-world scenarios involve additional factors:

- Heat Loss: In practical situations, some heat may be lost to the surroundings, resulting in a lower actual temperature increase.
- Phase Changes: If the substance undergoes a phase change (melting, boiling), additional heat (latent heat) is required without temperature change.
- Non-uniform Heating: Heat might not distribute evenly, leading to temperature gradients within the sample.

In controlled laboratory settings, calorimeters are used to minimize heat loss and ensure accurate measurements.

Conclusion and Practical Applications

This problem exemplifies fundamental principles of thermodynamics and calorimetry:

- Understanding how heat energy affects temperature based on specific heat capacity.
- Applying the heat transfer equation $Q = mc\Delta T$ to solve for unknowns.
- Recognizing the importance of material properties in thermal processes.

Applications include:

- Designing heating and cooling systems.
- Calculating energy requirements in industrial processes.
- Understanding environmental temperature regulation.
- Analyzing chemical reactions involving heat exchange.

In summary:

- The temperature increase of a substance upon heat absorption depends on its mass, specific heat capacity, and the amount of heat added.
- Accurate knowledge of the substance's properties is essential for precise calculations.
- The example demonstrates that, assuming water's specific heat capacity, the temperature of a 42.65 g sample initially at 25.1°C would rise to approximately 36.26°C after absorbing 476 calories.

By mastering these concepts, scientists and engineers can predict thermal behavior in various systems, optimize energy usage, and develop more efficient thermal management strategies.

Keywords: heat transfer, specific heat capacity, calorimetry, temperature change, thermal energy, calorimeter, heat absorption, thermal properties, thermodynamics, energy calculations

Frequently Asked Questions

What is the initial temperature of the substance before heating?

The initial temperature of the substance is 25.1°C.

How much heat energy is absorbed by the substance?

The substance absorbs 476 calories of heat energy.

What is the mass of the substance involved in the heating process?

The mass of the substance is 42.65 grams.

How do you calculate the change in temperature after heat absorption?

The change in temperature (ΔT) can be found using the formula $\Delta T = Q / (m \times c)$, where Q is heat absorbed, m is mass, and c is specific heat capacity.

What additional information is needed to find the specific heat capacity of the substance?

You need to know the temperature change (ΔT) to calculate the specific heat capacity using the formula $c = Q / (m \times \Delta T)$.

If the specific heat capacity of the substance is known, how can you determine its final temperature?

The final temperature can be calculated as $T_{\text{final}} = T_{\text{initial}} + (Q / (m \times c))$.

Why is understanding heat transfer important in material science?

Understanding heat transfer helps in determining a material's thermal properties, designing thermal systems, and enhancing energy efficiency in various applications.

Additional Resources

A 42.65 G Sample Of A Substance Is Initially At 25.1°C. After Absorbing 476 Cal Of Heat, The Temperature is a classic problem in thermodynamics and calorimetry, illustrating the fundamental principles of heat transfer, specific heat capacity, and phase transitions. This scenario offers a practical glimpse into how energy exchange affects the temperature of a substance, providing insights that are essential for chemists, physicists, and engineers alike. By dissecting this problem comprehensively, we can deepen our understanding of the underlying physics and chemistry, as well as appreciate the broader applications in industry, research, and daily life.

Understanding the Basic Concepts

Before delving into calculations and analysis, it's essential to clarify the core concepts that underpin this problem. These include the nature of heat transfer, specific heat capacity, and the relationship between heat, mass, temperature change, and specific heat.

Heat Transfer and Energy Exchange

Heat transfer is the process by which thermal energy moves from one system or object to another due to temperature differences. In this context, the substance absorbs heat from an external source, leading to an increase in its internal energy and, consequently, its temperature.

The unit of heat in this problem is the calorie (Cal), which is a common unit in calorimetry. Notably, 1 Calorie (kcal) equals 1,000 small calories (cal). The energy absorbed by the substance, 476 cal, will be used to calculate the resultant temperature change.

Specific Heat Capacity

The specific heat capacity (often simply called specific heat) is a material property indicating how much heat energy is needed to raise the temperature of one gram of a substance by one degree Celsius (or Kelvin). It is denoted as c and measured in units of $\text{cal}/(\text{g}\cdot^{\circ}\text{C})$.

Mathematically, the relationship between heat transfer, mass, specific heat, and temperature change is described by:

$$Q = mc\Delta T$$

Where:

- Q = heat absorbed or released (in cal)
- m = mass of the substance (in g)

- c = specific heat capacity (in $\text{cal}/(\text{g}\cdot^\circ\text{C})$)
- ΔT = change in temperature (in $^\circ\text{C}$)

This fundamental equation allows us to determine unknown quantities when the others are known.

Analyzing the Problem Step-by-Step

The problem provides:

- Initial mass $m = 42.65\text{ g}$
- Initial temperature $T_i = 25.1^\circ\text{C}$
- Heat absorbed $Q = 476\text{ cal}$
- Final temperature $T_f = ?$

Our goal is to find the final temperature after heat absorption, assuming no phase change or other energy-consuming processes.

Step 1: Recognize the Unknowns and Assumptions

Since the problem asks for the final temperature and provides the initial temperature, mass, and heat absorbed, the key unknown is T_f .

An important assumption here is that the specific heat capacity c remains constant over the temperature range, which is generally valid for small temperature changes and simple substances. If the substance undergoes phase changes or if c varies significantly with temperature, the problem becomes more complex.

Step 2: Determine the Specific Heat Capacity or Make Necessary Assumptions

If the specific heat capacity c of the substance is known, the calculation becomes straightforward:

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

However, if c is not provided, the problem cannot be solved definitively unless the substance is specified, or an average specific heat is assumed based on typical values.

Suppose the substance is water, which is common in calorimetry problems, with:

$$c_{\text{water}} = 1\text{ cal}/(\text{g}\cdot^\circ\text{C})$$

\]

Using this assumption, we can proceed to estimate the temperature change.

Calculating the Temperature Change

Step 3: Apply the Heat Equation

Using the heat transfer formula:

$$\begin{aligned} & \backslash[\\ Q &= mc\Delta T \\ & \backslash] \end{aligned}$$

Rearranged to solve for (ΔT) :

$$\begin{aligned} & \backslash[\\ \Delta T &= \frac{Q}{mc} \\ & \backslash] \end{aligned}$$

Substituting known values for water:

$$\begin{aligned} & \backslash[\\ \Delta T &= \frac{476 \text{ cal}}{42.65 \text{ g} \times 1 \text{ cal/(g}\cdot\text{°C)}} = \frac{476}{42.65} \\ & \approx 11.16^\circ\text{C} \\ & \backslash] \end{aligned}$$

This indicates that, assuming a specific heat of water, the temperature of the sample would increase by approximately 11.16°C.

Step 4: Determine the Final Temperature

Adding this change to the initial temperature:

$$\begin{aligned} & \backslash[\\ T_f &= T_i + \Delta T = 25.1^\circ\text{C} + 11.16^\circ\text{C} \approx 36.26^\circ\text{C} \\ & \backslash] \end{aligned}$$

Therefore, if the substance is water, the final temperature would be approximately 36.3°C after absorbing 476 cal of heat.

Critical Evaluation and Broader Implications

While the above calculation provides a straightforward estimate, it's crucial to consider the context and limitations.

1. Validity of Assumptions

- Substance Identification: The specific heat capacity used (water's) is common in such problems but might not apply to other substances. Different materials have vastly different specific heats; for metals, for example, values are much lower (around $0.1 \text{ cal/(g}\cdot\text{°C)}$), whereas for substances like aluminum ($\sim 0.215 \text{ cal/(g}\cdot\text{°C)}$) or glass ($\sim 0.2 \text{ cal/(g}\cdot\text{°C)}$), the value varies.
- Constant Specific Heat: The assumption that c remains constant over the temperature range is an approximation. Many substances exhibit temperature-dependent specific heats, especially over larger temperature spans.
- No Phase Changes: The calculation assumes the substance remains in the same phase (liquid, solid, or gas). If the temperature approached a melting or boiling point, additional energy would go into phase transition rather than temperature increase, complicating the analysis.

2. Energy Required for Phase Changes and Latent Heat

If the heat input were sufficient to cause melting or vaporization, the energy would be partitioned into changing phases rather than raising temperature. For example, water requires:

- 80 cal/g to melt from ice at 0°C
- 540 cal/g to vaporize at 100°C

In such cases, the temperature would plateau during phase transitions until completion, and the calculation would need to incorporate latent heats.

3. Practical Applications and Significance

Understanding the thermal behavior of substances is fundamental in:

- Designing thermal systems: Heat exchangers, reactors, and insulation materials depend on precise calculations of temperature changes.
- Food Industry: Cooking times and temperatures are based on heat transfer principles.
- Pharmaceuticals: Stability of compounds often hinges on temperature control during processing.
- Environmental Science: Climate models incorporate heat capacity of water bodies and land.

Advanced Considerations and Real-World Complexities

In real-world scenarios, several additional factors influence heat transfer and temperature changes.

1. Heat Loss and Environmental Factors

- Conduction, Convection, and Radiation: In experimental setups, heat loss to surroundings can significantly affect the temperature change.
- Insulation: Proper insulation minimizes heat loss, ensuring that absorbed heat results primarily in temperature change of the sample.

2. Non-Idealities in Materials

- Impurities: Can affect specific heat capacity.
- Structural Changes: Heating may lead to structural modifications, affecting thermal properties.
- Container Effects: The material and shape of the container can influence heat transfer efficiency.

3. Measurement Precision

- Temperature Readings: Accurate thermometry is crucial; calibration errors can lead to discrepancies.
- Heat Measurement: Calorimeters must be properly calibrated to ensure the heat input is accurately measured.

Conclusion: Integrating Theory and Practice

The problem of a 42.65 g sample absorbing 476 cal and resulting in a temperature increase exemplifies fundamental thermodynamic principles. When assuming the substance as water, the calculations suggest an approximate temperature rise of 11.16°C , culminating in a final temperature of around 36.3°C . This example underscores the importance of understanding specific heat capacity and the assumptions involved in such calculations.

However, real-world applications demand careful consideration of additional factors like phase changes, heat losses, material properties, and measurement accuracy. The principles illustrated here serve as a foundation for more complex analyses in scientific research, engineering design, and industrial processes. Mastery of these concepts enables professionals to optimize thermal systems, predict material behavior under heating, and develop innovative solutions across diverse fields.

In essence, this problem is not just about a numerical answer but about appreciating the intricate interplay of energy, matter, and temperature—a dance governed by the laws of thermodynamics that

shape our daily lives and technological advancements alike.

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