If A System Does 84.0 Kj Of Work On Its Surroundings And Releases 105 Kj Of Heat, What Is The Change

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Understanding the thermodynamic behavior of systems and their surroundings is fundamental in physics and chemistry. When analyzing such processes, it is crucial to determine the change in the system's internal energy, often denoted as ΔU. This involves considering the work done by or on the system, heat exchange, and the principles governing energy conservation. In this article, we will explore how to calculate the change in internal energy when a system performs work and releases heat, exemplified by the given data: the system does 84.0 kJ of work and releases 105 kJ of heat.

Fundamental Concepts in Thermodynamics

Before delving into the specific problem, it is essential to review some core thermodynamic principles and terminology.

Internal Energy (U)

- The total energy contained within a system, including kinetic and potential energies of particles.
- Changes in internal energy (ΔU) are central to understanding energy transformations during processes.

Work (W)

- Energy transfer resulting from macroscopic forces acting over a distance.
- When a system does work, it transfers energy to its surroundings.
- The sign convention varies; here, we will assume that work done by the system on surroundings is positive, which is standard in many thermodynamics contexts.

Heat (Q)

- Energy transfer due to temperature difference.
- Releasing heat (exothermic process) implies the system's temperature decreases or

energy leaves the system.

First Law of Thermodynamics

This law states that:

$$
\Delta U = Q + W
$$

- If Q is positive, the system absorbs heat.
- If Q is negative, the system releases heat.
- If W is positive, the system does work on surroundings.

Note: The sign conventions are essential; in many textbooks, work done by the system is considered positive, and heat added to the system is positive.

Analyzing the Given Data

The problem provides:

- Work done by the system: 84.0 kJ
- Heat released by the system: 105 kJ

Using the first law:

$$
\Delta U = Q + W
$$

But first, clarify the sign conventions:

- The system does work on surroundings: $W = +84.0\, \text{kJ}$
- The system releases heat: since heat is released, $Q = -105\, \text{kJ}$

Substituting into the equation:

$$
\Delta U = (-105\, \text{kJ}) + (84.0\, \text{kJ}) = -21.0\, \text{kJ}
$$

Result: The change in the system's internal energy is $-21.0\, \text{kJ}$.

This negative value indicates that the system has lost energy during the process.

Step-by-Step Calculation Process

To ensure clarity, let’s walk through the precise steps involved in determining the internal energy change:

Step 1: Identify the signs of heat and work

- Work done by the system: +84.0 kJ
- Heat released by the system: -105 kJ

Step 2: Write the first law equation

$$
\Delta U = Q + W
$$

where:

- $Q = -105\, \text{kJ}$
- $W = +84.0\, \text{kJ}$

Step 3: Plug in the values

$$
\Delta U = -105\, \text{kJ} + 84.0\, \text{kJ} = -21.0\, \text{kJ}
$$

Step 4: Interpret the result

- The negative ΔU indicates the system's internal energy decreases.
- The magnitude (21.0 kJ) shows the amount of energy lost from the system.

Implications of the Calculated Change

Understanding the change in internal energy helps in predicting the behavior of the system:

1. **Energy Loss:** The system releases more heat than the work it performs, resulting in an overall energy decrease.

2. **Thermodynamic State:** The process is exothermic, as energy is leaving the system in the form of heat.

3. **Practical Applications:** Such calculations are essential in designing engines, refrigerators, and chemical reactors where energy transfer is critical.

Real-World Examples and Applications

Understanding energy changes in systems is vital across various scientific and engineering fields. Here are some practical examples:

1. Combustion Engines

- Fuel combustion releases heat and performs work (e.g., moving pistons).
- Calculating energy changes helps optimize efficiency.

2. Chemical Reactions

- Exothermic reactions release heat.
- Monitoring energy changes aids in safety and process control.

3. Refrigeration and Air Conditioning

- Work input and heat transfer calculations determine system efficiency.
- Understanding energy flow ensures optimal performance.

Key Takeaways for Students and Professionals

- Always pay attention to sign conventions in thermodynamic calculations.
- The first law of thermodynamics provides a straightforward way to evaluate energy changes.
- Real-world systems often involve a combination of work and heat transfer, requiring careful analysis.

Conclusion

In summary, when a system does 84.0 kJ of work on its surroundings and releases 105 kJ of heat, the change in its internal energy is \(-21.0\) kJ. This negative change signifies an overall energy loss, characteristic of an exothermic process where more heat leaves the system than work is performed. Mastery of these concepts and calculations is fundamental for professionals working in thermodynamics, energy management, chemical engineering, and related fields. Accurate evaluation of energy changes enables better system design, efficiency optimization, and safety assurance.

Meta Description: Learn how to calculate the change in a system's internal energy when it performs work and releases heat, with detailed explanations and step-by-step guidance based on thermodynamics principles.

Frequently Asked Questions

What is the change in the internal energy of the system if it does 84.0 kJ of work on its surroundings and releases 105 kJ of heat?

The change in internal energy (ΔU) is -189 kJ, calculated as $\Delta U = Q - W = (-105 \text{ kJ}) - (84 \text{ kJ}) = -189$ kJ.

How is the first law of thermodynamics applied to this problem?

The first law states $\Delta U = Q - W$. Since the system releases heat ($Q = -105$ kJ) and does work ($W = 84.0$ kJ), the change in internal energy is $\Delta U = -105 \text{ kJ} - 84.0 \text{ kJ} = -189$ kJ.

What does a negative ΔU indicate about the system's energy?

A negative ΔU indicates that the system's internal energy decreases during the process.

If the system performs work and releases heat, is it exothermic or endothermic?

It is exothermic, as the system releases heat to its surroundings.

How do you calculate the change in internal energy for this process?

Use the formula $\Delta U = Q - W$. Given $Q = -105$ kJ and $W = 84.0$ kJ, $\Delta U = -105 \text{ kJ} - 84.0 \text{ kJ} = -189$ kJ.

What is the significance of the work done by the system in this scenario?

The work done by the system on its surroundings contributes to the decrease in its internal energy, indicating energy transfer out of the system.

Can we say the process is spontaneous based solely on the energy change?

No, the spontaneity depends on additional factors like entropy; energy change alone does not determine spontaneity.

Is the system gaining or losing energy from the surroundings?

The system is losing energy, as it releases heat and performs work, resulting in a decrease in internal energy.

What would be the change in internal energy if the heat released was 50 kJ instead of 105 kJ, with the same work done?

The change would be $\Delta U = -50 \text{ kJ} - 84.0 \text{ kJ} = -134$ kJ.

How does this problem illustrate the conservation of energy in thermodynamics?

It demonstrates that energy lost as heat and work results in a decrease in the system's internal energy, consistent with energy conservation principles.

Additional Resources

Understanding the Thermodynamic Change of the System: Work, Heat, and Internal Energy

In the realm of thermodynamics, analyzing how a system's energy changes during processes involving work and heat transfer is fundamental. The problem in question involves a system that performs 84.0 kJ of work on its surroundings and releases 105 kJ of heat. To understand the implications of these values, we need to delve deeply into the concepts of work, heat, and the first law of thermodynamics, as well as interpret what the change in the system's internal energy signifies.

Fundamental Concepts in Thermodynamics

Before tackling the specific problem, it is essential to clarify some key concepts:

1. Internal Energy (U)

- The total energy contained within a system, including kinetic, potential, chemical, and other forms of energy.
- Changes in internal energy (ΔU) are central to understanding thermodynamic processes.

2. Work (W)

- Energy transferred when a force causes displacement.
- In thermodynamics, work often takes the form of boundary work, such as expansion or compression of gases.
- Conventionally, work done by the system on the surroundings is considered positive, whereas work done on the system is negative.

3. Heat (Q)

- Energy transferred due to temperature difference.
- Heat absorbed by the system is positive; heat released (or rejected) by the system is negative.

4. The First Law of Thermodynamics

- The law states that the change in internal energy of a system is equal to the heat added to the system minus the work done by the system:

$$
\Delta U = Q - W
$$

- Note: When applying the first law, it is critical to maintain consistent sign conventions.

Sign Conventions and Their Impact on Interpretation

Different textbooks and disciplines adopt different sign conventions. The most common conventions are:

- Physics convention:
- $Q > 0$ when heat is added to the system.
- $W > 0$ when work is done by the system.
- Chemistry/Engineering convention:
- $Q > 0$ when heat is added to the system.
- $W > 0$ when work is done on the system.

In this problem, the statement "performs 84.0 kJ of work on its surroundings" indicates that the system is doing work by itself, so:

- $W = 84.0\, \text{kJ}$ (positive if we follow the convention that work done by the system is positive).

The heat release of 105 kJ indicates the system is releasing heat, so:

- $Q = -105\, \text{kJ}$ (negative because heat is leaving the system).

This interpretation aligns with the thermodynamic sign conventions used in physics.

Calculating the Change in Internal Energy (ΔU)

Using the first law:

$$
\Delta U = Q - W
$$

Substituting the known values:

$$
\Delta U = (-105\, \text{kJ}) - (84.0\, \text{kJ}) = -189\, \text{kJ}
$$

Interpretation:

- The negative sign indicates that the internal energy of the system decreases by 189 kJ during the process.

- The system has lost energy overall, primarily due to heat loss and work done on the surroundings.

Physical Interpretation of the Energy Change

What does a ΔU of -189 kJ signify?

- The system has undergone a process where it loses energy, reducing its internal energy.
- The energy loss manifests as:
- Releasing heat to the surroundings.
- Performing work on the surroundings.

Implications:

- The process could be akin to cooling combined with expansion.
- Such a process might occur in real-world applications like engine exhaust, chemical reactions releasing energy, or cooling systems performing work.

Deeper Insights into the Process

To understand the process thoroughly, consider the following aspects:

1. Energy Flow Summary

Aspect	Quantity	Significance
Heat released	105 kJ	The system loses heat to surroundings
Work done	84.0 kJ	The system does work on surroundings
Internal energy change	-189 kJ	Overall energy decrease

The total energy leaving the system (heat + work) sums to 189 kJ, matching the decrease in internal energy.

2. Nature of the Process

- The system is exothermic, since it releases heat.
- It performs work on its surroundings, possibly expanding or causing other energy transfer mechanisms.
- The process might resemble a cooling expansion or a chemical process where energy is

released as heat and work.

3. Possible Real-World Examples

- An idealized engine cycle where energy is expelled as heat, and work is done during an expansion.
- A chemical reaction releasing heat and performing work, such as an exothermic combustion process.
- A cooling device where energy is removed, and mechanical work is performed.

Broader Significance and Applications

Understanding how energy quantities like work and heat influence the internal energy of a system is crucial across various fields:

1. Engineering Applications

- Designing engines and turbines that optimize work output and heat rejection.
- Thermodynamic cycle analysis for power plants, refrigerators, and air conditioners.

2. Chemical Processes

- Calculating energy changes in reactions for safety, efficiency, and environmental impact.
- Understanding exothermic reactions and their energy budgets.

3. Environmental Considerations

- Assessing energy efficiency and waste heat management.
- Modeling natural processes like geothermal energy transfer.

Summary and Key Takeaways

- The change in the system's internal energy during the process is $\boxed{\Delta U = -189\, \text{kJ}}$.
- The negative sign indicates a net loss of energy from the system, consistent with heat being released and work being performed.
- Sign conventions are critical; in this context, the system performs work and releases heat, both leading to a decrease in internal energy.

- Real-world processes involve similar energy transfers, emphasizing the importance of thermodynamic analysis in engineering, chemistry, and environmental science.

Final Reflection: Significance of Thermodynamic Analysis

This problem exemplifies the importance of understanding the first law of thermodynamics and the sign conventions used to interpret energy transfer processes. It also highlights how quantifying work and heat allows engineers and scientists to analyze system efficiencies, predict behavior, and design better energy systems. Whether dealing with engines, chemical reactions, or natural phenomena, grasping the interplay between work, heat, and internal energy is fundamental to advancing technology and understanding the physical world.

In conclusion, performing a detailed analysis of energy transfers reveals that the system’s internal energy decreases by 189 kJ during the process characterized by work done on surroundings and heat released. This comprehensive understanding underscores the importance of precise sign conventions and the fundamental principles governing thermodynamic processes.

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