

Enthalpy Changes and Calorimetry

Time: 20 min. 6 questions.

Instructions

Answer all six questions. For multiple choice, select the best answer. For short answer, write a complete sentence or brief explanation. For true or false, state whether the statement is correct. For the numeric question, show your calculation and include units in your final answer. You may use a calculator.

Questions

1. A student burns 2.5 g of ethanol in a bomb calorimeter. The temperature of 1000 g of water surrounding the bomb rises by 8.2 °C. Calculate the heat released by the combustion, using $Q = mc\Delta T$ (Use $c = 4.18 \text{ J/g}\cdot^\circ\text{C}$ for water.)
([Outcome](#))
2. Which of the following statements best describes enthalpy?
A) The total kinetic energy of all particles in a system
B) The heat content of a substance at constant pressure
C) The energy required to break all bonds in a molecule
D) The rate at which a reaction proceeds
([Outcome](#))
3. A reaction has $\Delta H^\circ = -125 \text{ kJ/mol}$. Explain what this notation tells you about the reaction and whether it is endothermic or exothermic.
([Outcome](#))
4. True or false: An endothermic reaction absorbs heat from the surroundings, resulting in a negative ΔH° value.
([Outcome](#))
5. Using standard enthalpies of formation, calculate ΔH° for the reaction $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$, given that $\Delta H^\circ_f[\text{H}_2\text{O}(\text{l})] = -285 \text{ kJ/mol}$, and $\Delta H^\circ_f[\text{H}_2(\text{g})] = 0 \text{ kJ/mol}$, and $\Delta H^\circ_f[\text{O}_2(\text{g})] = 0 \text{ kJ/mol}$.
([Outcome](#))
6. A student performs a calorimetry experiment to measure the enthalpy of combustion of a candle. Explain why the measured value is typically lower than the theoretical value found in reference tables.
([Outcome](#))

Curriculum Coverage

Question	Outcome Alias	Specific Outcome
1	Outcome	Apply $Q = mc\Delta T$ to transfer analysis
2	Outcome	Define enthalpy and molar enthalpy
3	Outcome	Use and interpret ΔH° , ΔH°_f
4	Outcome	Classify reactions as endothermic or exothermic
5	Outcome	Predict enthalpy change using standard enthalpies of formation

6	Outcome	Use calorimetry data to determine enthalpy changes
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Answer Key

1. $Q = (1000 \text{ g})(4.18 \text{ J/g} \cdot ^\circ\text{C})(8.2 ^\circ\text{C}) = 34,276 \text{ J}$ "H 34.3 kJ; this represents the heat absorbed by the water and released by the combustion.
2. **B) The heat content of a substance at constant pressure.** Enthalpy is defined as the heat content of a system at constant pressure and is a state function used to describe energy changes in reactions.
3. **The negative sign indicates that 156 kJ of energy is released per mole of reaction; this is an exothermic reaction.** Negative ΔH values always correspond to exothermic processes in which heat flows out of the system.
4. **False.** Endothermic reactions absorb heat from the surroundings, which results in a positive ΔH value, not negative.
5. **572 kJ** Using $\Delta H_{\text{rxn}} = \sum \Delta H_f^\circ(\text{products}) - \sum \Delta H_f^\circ(\text{reactants})$ the reaction is highly exothermic.
6. **Heat is lost to the surroundings during the experiment through the calorimeter walls and to the air, so less heat is captured than the theoretical value predicts.** Calorimeters are not perfectly insulated, and incomplete combustion or other experimental losses also reduce the measured enthalpy.