

Enthalpy Changes and Calorimetry

30 minutes, 3 stages

Calorimetry is how chemists measure energy changes they cannot see directly. You will define enthalpy and molar enthalpy ([Outcome](#), [Outcome](#)), apply $Q = mc\Delta T$ and $Q = n\Delta H$ ([Outcome](#), [Outcome](#)), and use calorimetry data to calculate the enthalpy change of a reaction ([Outcome](#)). These skills also require correct sign conventions ([Outcome](#)) and connect to real calorimeter design, such as the polystyrene-cup calorimeter ([Resource](#)) and the nut-combustion lab ([Resource](#)).

Stage 1: Warm-up (8 min)

Answer these in complete sentences.

1. Define enthalpy in your own words. State what a negative ΔH indicates about a reaction ([Outcome](#), [Outcome](#)).
2. Define molar enthalpy and explain how it differs from enthalpy for a single reaction ([Outcome](#)).
3. Write the equation $Q = mc\Delta T$ and label each variable with its symbol, meaning, and SI units ([Outcome](#)).
4. A calorimetry experiment loses some heat to the surroundings, so the water in the calorimeter does not receive all the energy released by the reaction. Explain in one or two sentences why this causes the calculated enthalpy change to be lower in magnitude than the true value ([Resource](#)).

Stage 2: Worked calculations (16 min)

Work through the following problems. Show all substitutions and units. Use $c(\text{water}) = 4.18 \text{ J/(g}\cdot^\circ\text{C)}$ unless a problem states otherwise.

Problem 1. A student burns a sample of a snack food (a common calorimetry demonstration, [Resource](#)) under a calorimeter containing 150.0 g of water. The water temperature rises from 21.4°C to 63.8°C.

- a) Calculate Q , the heat absorbed by the water, in joules.
- b) Convert your answer to kilojoules.

Problem 2. In a coffee-cup calorimeter, 50.0 g of water at 22.0°C is mixed with 50.0 g of water at 80.0°C. Assume no heat is lost to the surroundings and the final temperature is 51.0°C.

- a) Calculate the heat lost by the hot water using $Q = mc\Delta T$.
- b) Calculate the heat gained by the cold water using $Q = mc\Delta T$.
- c) Compare your two answers from (a) and (b). Explain what this comparison tells you about whether your calorimeter behaved as an ideal, heat-tight system ([Outcome](#), [Resource](#)).

Problem 3. A student reacts 0.0500 mol of a substance in 100.0 g of water inside a calorimeter. The water temperature increases by 8.20°C.

- a) Calculate Q for the water, in joules, then convert to kilojoules ([Outcome](#)).
- b) Calculate the molar enthalpy of the reaction in kJ/mol, using the number of moles reacted ([Outcome](#), [Outcome](#)).
- c) Write your final answer using correct sign convention to indicate whether the reaction is endothermic or exothermic ([Outcome](#)).

Stage 3: Consolidation (6 min)

1. In two or three sentences, explain why calorimetry data (measured Q for water) can be used to calculate ΔH for a chemical reaction, even though the calorimeter measures the water's temperature change and not the reaction directly ([Outcome](#)).

2. List two design features of a calorimeter, such as the polystyrene-cup design ([Resource](#)), that reduce heat loss to the surroundings and therefore improve the accuracy of an experimental enthalpy value.

Self-check

1. If your calculated $q_{\text{calorimeter}}$ for a calorimetry experiment has a smaller magnitude than the accepted value, what is the most likely experimental cause?
2. A reaction releases 12.6 kJ of heat when 0.0300 mol reacts. What is the molar enthalpy, and what sign should you assign to $q_{\text{calorimeter}}$ for this exothermic process?
3. Explain, in your own words, the difference between Q in $Q = mc\Delta T$ and ΔH for the reaction that produced it.