

Enthalpy Changes and Calorimetry

Using measured temperature changes to calculate the energy released or absorbed in a chemical reaction.



Figure 1: A polystyrene-cup calorimeter setup on a lab bench, showing a thermometer inserted through a cover into a cup of water, a small burning

Learning intentions

We are learning to calculate heat transfer using $Q = mc\Delta T$ and calorimetry data to determine the enthalpy change of a reaction, expressing that change with correct sign.

Success criteria

I can define enthalpy and molar enthalpy for a chemical reaction.

I can apply $Q = mc\Delta T$ and $Q = n\Delta H$ to calculate the heat transfer for a substance.

I can use calorimetry data (mass, specific heat capacity, temperature change) to calculate q , for a reaction, including molar enthalpy.

I can write and interpret q , including sign conventions for exothermic and endothermic processes.

Curriculum alignment

- 30-A1.1k: recall the application of $Q = mc\Delta T$ [Outcome](#)
- 30-A1.3k: define enthalpy and molar enthalpy for chemical reactions ([Outcome](#), [Outcome](#))
- 30-A1.5k: use and interpret q , ΔH and calculate energy changes in chemical reactions ([Outcome](#))
- 30-A1.8k: use calorimetry data to determine the enthalpy changes in chemical reactions ([Outcome](#))

Materials

- Whiteboard or projector for worked examples
- Student calculators
- Handout with the calorimetry data table (mass of water, initial and final temperatures, mass of fuel burned) for the guided and independent practice problems
- Specific heat capacity of water reference value (4.19 J/(g·°C))
- Optional: photographs or short clip from the nut-combustion calorimetry investigation for context ([Resource](#))
- Optional: reference to the polystyrene-cup calorimeter construction lab for a follow-up hands-on lesson ([Resource](#))

Lesson sequence

1. Hook (5 minutes)

Write on the board: "A 2.10 g sample of a peanut is burned under a can of water. The water's temperature rises from 22.0°C to 45.5°C. The can holds 100.0 g of water."

Ask the class: "Without doing any math yet, which direction is the energy moving, from the peanut to the water, or from the water to the peanut? How do you know?"

Take two or three verbal answers. Confirm: energy moves from the burning peanut to the water, so the water gains heat and the reaction (combustion) releases heat. This is the setup used in [Resource](#).

2. Direct instruction (10 minutes)

State the definitions: enthalpy (H) is the heat content of a system at constant pressure. Enthalpy change (q), for a reaction at constant pressure. Molar enthalpy is the enthalpy change per mole of a specified substance in the reaction, expressed in kJ/mol ([Outcome](#)).

Write the sign convention on the board:

Exothermic reaction: heat is released, q , ΔT , temperature of the surroundings (the water) increases.

Endothermic reaction: heat is absorbed, q , ΔT , temperature of the surroundings decreases.

Present the formula: $Q = mc\Delta T$, m is the mass of the substance absorbing or releasing heat (usually the water) in grams, c is the specific heat capacity in J/(g·°C), and ΔT is the final temperature minus the initial temperature ([Outcome](#)).

Work the hook example on the board step by step:

$$Q = mc\Delta T$$

$$Q = (100.0 \text{ g})(4.19 \text{ J/(g}\cdot^\circ\text{C)})(45.5^\circ\text{C} - 22.0^\circ\text{C})$$

$$Q = (100.0 \text{ g})(4.19 \text{ J/(g}\cdot^\circ\text{C)})(23.5^\circ\text{C})$$

$$Q = 9846.5 \text{ J}$$

$$Q = 9.85 \text{ kJ}$$

State: "This is the heat gained by the water. Since energy is conserved, the peanut released 9.85 kJ. Because the reaction released energy, the enthalpy change for the reaction is negative: ΔH for this sample" ([Outcome](#), [Outcome](#)).

Show how to convert to molar enthalpy. Molar mass of the fuel is needed. If the peanut sample has an approximate molar mass and 2.10 g corresponds to 0.0120 mol of fuel (fictional value for illustration), then:

$$\Delta H_{\text{molar}} = \Delta H \div n$$

$$\Delta H_{\text{molar}} = \Delta H \div \left(\frac{\text{mass}}{\text{molar mass}} \right)$$

$$\Delta H_{\text{molar}} = \Delta H \div \left(\frac{2.10 \text{ g}}{175 \text{ g/mol}} \right)$$

3. Guided practice (10 minutes)

Distribute or display this problem and work through it together, calling on students for each step:

"In a calorimetry trial, 150.0 g of water is heated from 19.5°C to 38.2°C by burning 1.35 g of a fuel with a molar mass of 46.07 g/mol (ethanol). Calculate the heat absorbed by the water, the enthalpy change for the combustion of this sample, and the molar enthalpy of combustion."

Ask: "What is q for the water?" $q = mc\Delta T$

Ask: "What is Q ?" Have a student calculate at the board:

$$Q = (150.0 \text{ g})(4.19 \text{ J/(g}\cdot^\circ\text{C)})(18.7^\circ\text{C}) = 11,753 \text{ J} = 11.75 \text{ kJ}$$

Ask: "What is ΔH for the reaction, and what sign does it have, and why?" (ΔH for the reaction, negative because combustion is exothermic and the water gained heat.)

Ask: "How many moles of ethanol were burned?"

$$n = 1.35 \text{ g} \div 46.07 \text{ g/mol} = 0.0293 \text{ mol}$$

Ask: "What is the molar enthalpy of combustion?"

$$\Delta H_{\text{molar}} = \Delta H \div n = 11.75 \text{ kJ} \div 0.0293 \text{ mol} = -401 \text{ kJ/mol}$$

Checkpoint question to the whole class: "Why do we divide by moles instead of grams to get molar enthalpy?" Listen for: molar enthalpy compares reactions on a per-particle (per-mole) basis rather than a per-gram basis, which allows fair comparison between different substances.

4. Independent practice (10 minutes)

Students work individually or in pairs on this problem, writing full solutions with units at each step:

"A student burns 0.85 g of a candle wax sample under 200.0 g of water. The water temperature rises from 21.0°C to 33.4°C. The molar mass of the wax is 226 g/mol. Calculate: (a) the heat gained by the water, (b) the enthalpy change for the sample burned, with correct sign and ΔH for the reaction, (c) the molar enthalpy of combustion for the wax."

Circulate and check that students are:

using $q = mc\Delta T$ with initial temperature.

carrying the correct sign through from Q to ΔH ,

converting grams of fuel to moles before dividing.

5. Consolidation (5 minutes)

Ask the whole class: "If this had been an endothermic process instead, for example an ice pack absorbing heat from your hand, what would change in this calculation?" Listen for: the water (or surroundings) would lose heat, q_{GB} could be negative, and q_{D} , or the process would be reported as positive.

Differentiation

Support: Provide a filled-in formula triangle or step template ($Q = mc\Delta T$ with blanks for each variable, followed by a separate line for $n = \text{mass} \div \text{molar mass}$, followed by $Q = n\Delta H$) so students can substitute values without holding the full sequence in memory. Reduce the independent practice problem to two steps (calculate Q , then state $Q = n\Delta H$) over the molar enthalpy step as a stretch goal rather than a requirement.

Assessment

Look for: correct substitution into $Q = mc\Delta T$, correct calculation of Q (not the reverse), the correct sign assigned to Q , & 6Vb $q = n\Delta H$ F+R pater gained or lost heat, and correct conversion from grams to moles before calculating molar enthalpy. A student who reaches a numerically correct Q but assigns the wrong sign to Q , ΔH V-0k one-on-one reminder on the exothermic and endothermic convention before the next lesson.