

The Calorimetry Puzzle

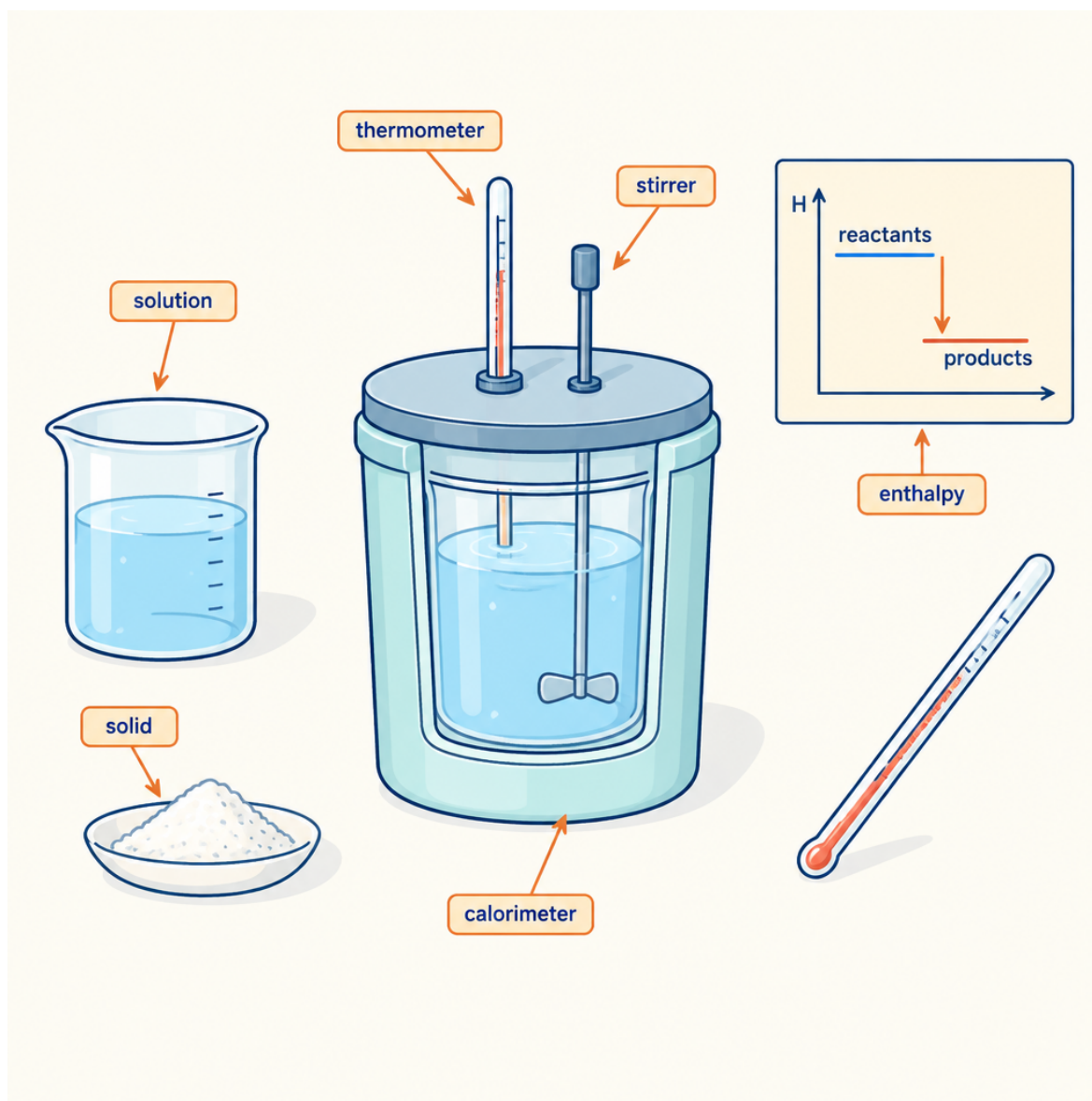


Figure 1: Clean labelled diagram in the style of a textbook figure illustrating Enthalpy changes and calorimetry: the equipment, materials and key

This brain teaser applies the heat transfer equation $Q = mc\Delta T$ ([Outcome](#), [Outcome](#)) and the definition of enthalpy for a chemical reaction ([Outcome](#), [Outcome](#)) to a calorimetry data set, matching the way calorimetry data are used to determine enthalpy changes in the Chemistry 30 program ([Outcome](#)). The scenario mirrors the type of energy transfer measurement described in the nut-combustion calorimetry investigation ([Resource](#)) and connects to General Outcome 1, determining and interpreting energy changes in chemical reactions ([Topic](#)).

The puzzle

A student burns a small sample of fuel under a calorimeter containing 200.0 g of water. The water temperature rises from 21.5°C to 46.5°C. The specific heat capacity of water is 4.18 J/(g·°C).

The fuel sample has a molar mass of 46.0 g/mol, and the student burns 0.230 g of it during the trial.

Using $q = mc\Delta T$ calculate the heat absorbed by the water, q , using $q = mc\Delta T$. Remember that ΔT is the change in temperature, final minus initial.

Assume no heat is lost to the surroundings or absorbed by the calorimeter itself.

Hints

Hint 1: Start by finding the heat absorbed by the water, q , using $q = mc\Delta T$. Remember that ΔT is the change in temperature, final minus initial.

Hint 2: Once you have q in joules, convert it to kilojoules, then find the number of moles of fuel burned using the mass of fuel and its molar mass.

Hint 3: Divide the heat released (in kJ) by the number of moles of fuel burned to get the molar enthalpy. Since the water gained heat, the fuel must have released it, so the sign of ΔH , or the reaction should be negative.

Answer key

Step 1: Calculate q for the water.

$$q = mc\Delta T$$

Step 2: Calculate q , the heat absorbed by the water.

$$q = mc\Delta T$$

$$q = (200.0 \text{ g})(4.18 \text{ J/(g}\cdot\text{°C)})(25.0\text{°C})$$

$$q = 20900 \text{ J}$$

$$q = 20.9 \text{ kJ}$$

Step 3: Calculate the moles of fuel burned.

$$n = \text{mass} \div \text{molar mass}$$

$$n = 0.230 \text{ g} \div 46.0 \text{ g/mol}$$

$$n = 0.00500 \text{ mol}$$

Step 4: Calculate the molar enthalpy of combustion.

$$\Delta H = -q/n$$

$$\Delta H = -20.9 \text{ kJ} / 0.00500 \text{ mol}$$

$$\Delta H = -4180 \text{ kJ/mol}$$

The negative sign shows that the reaction releases energy to the surroundings, so the reaction is exothermic. The water absorbed the heat given off by the burning fuel, which is why the water's temperature rose while the fuel's chemical energy decreased.