

Measuring enthalpy change · supported focus

International Baccalaureate · IB Diploma · Chemistry · SL

Scope and prerequisites

Supported SL focus. First assessment 2025; current brief acquired; full Chemistry guide not acquired. Remaining guide, assessment and practical requirements retain their recorded holds.

Prerequisites: read the stated quantities and units, use arithmetic and the model conditions below. Each lesson develops its own method before independent transfer.

These are original or explicitly fictional teaching examples, not actual measurements or completed assessed learner investigations.

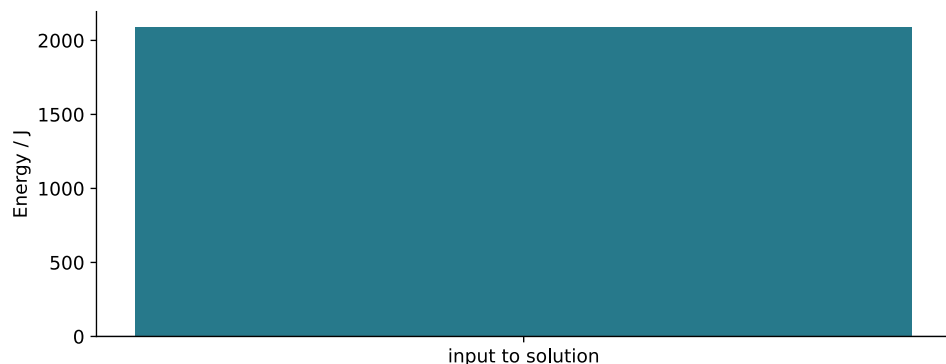
Calorimetry and chemical energy

■ What would explain this observation?

- A cup warms when two solutions react. The temperature rise measures energy transferred to the surroundings; it does not directly equal the **enthalpy change**.
- Start with a prediction. State the quantities or features you would compare, then decide what evidence could distinguish two explanations.

■ Build the model

- **Exothermic** reactions transfer energy to surroundings. Endothermic reactions take energy from surroundings. Bond breaking requires energy; bond formation releases energy.
- exothermic: Transferring energy to the surroundings; enthalpy change: Heat change at constant pressure for a stated process.



■ Choose evidence that can test it

- Use energy transferred = mass $\times$ specific heat capacity $\times$ temperature change. Convert joules to kilojoules before dividing by reaction amount. An exothermic molar enthalpy change has a negative sign.
- Use insulation and a lid, measure starting temperatures consistently, stir, and record a temperature-time series. Estimate the reaction temperature from an appropriate extrapolation rather than ignoring cooling during measurement.

■ Work from known quantities

- State the known values and their units. Choose the relation because its assumptions fit this case, then rearrange before substitution.
- Known: 100 g solution rises by 5.0 °C; specific heat capacity is 4.18 J per gram per degree. $q = mc\Delta T$. $q = 100 \times 4.18 \times 5.0 = 2,090 \text{ J} = 2.09 \text{ kJ}$. If 0.050 mol reacts, $\Delta H = -q/n = -2.09/0.050 = -41.8 \text{ kJ per mole}$.

Example:

50 g water rises 4 °C. Use $c = 4.2 \text{ J per gram per degree}$ to find q . Use the same sequence: known quantities $\rightarrow$ model $\rightarrow$ relation $\rightarrow$ substitution $\rightarrow$ unit and interpretation.

■ Check the conclusion and its limits

- Heat loss usually lowers the observed temperature rise. The solution gaining heat and the reaction losing heat have opposite signs.
- Return to the original observation. Explain what the result supports, which conditions it assumes, and one way to test a competing explanation.

Warn:

Bond breaking releases energy. This claim is false: Heat loss usually lowers the observed temperature rise. The solution gaining heat and the reaction losing heat have opposite signs.

Key:

Calorimetry and chemical energy: Use energy transferred = mass $\times$ specific heat capacity $\times$ temperature change. Convert joules to kilojoules before dividing by reaction amount. An exothermic molar enthalpy change has a negative sign.