

Measuring enthalpy change

International Baccalaureate · IB Diploma · Chemistry · SL

Scope and prerequisites

This handout teaches Reactivity 1.1, Measuring enthalpy changes, for the course first assessed in 2025. This unit has the same content at SL and HL; there are no additional higher-level understandings here.

Prerequisites: balancing equations, $n = m/M$, $n = cV$, limiting reactants and the particle model. All experimental data and questions are original, invented teaching examples.

Energy transfer: choose the system

A reacting mixture warms an insulated cup. Has energy been created? No: energy has been transferred from chemical stores into thermal motion. Total energy is conserved.

The **system** is the part studied; everything else is the **surroundings**. Here the chemical reaction is the system; the solution and cup are the thermal surroundings. The solution contains the reacting particles while also absorbing heat from their reaction.

Heat is energy transferred because of a temperature difference, measured in joules.

Temperature relates to average particle kinetic energy, not an amount of energy.

Two water samples at the same temperature can require different heat transfers for the same rise because their masses differ.

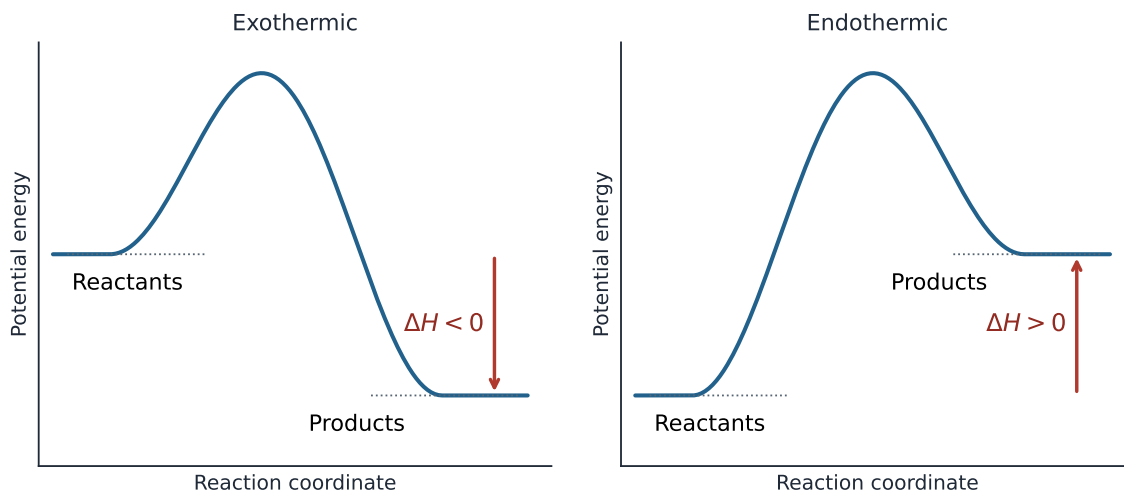
An **exothermic** reaction transfers energy from the system to the surroundings: the solution normally warms and the reaction has a negative **enthalpy change**.

An **endothermic** reaction takes energy from its surroundings: the solution normally cools and the reaction has a positive enthalpy change. A cooling solution has lost thermal energy even though the reaction has gained energy.

At constant pressure, with only pressure–volume work, the heat transferred to the reacting system equals its enthalpy change. We measure an enthalpy *difference*, not the absolute enthalpy of a substance.

Reaction profiles and relative stability

An energy profile plots **potential energy** against **reaction coordinate**: progress along a pathway, not time. Label both axes and the reactant and product levels; connect the levels through a peak.



The sign of the enthalpy change comes from the product level minus the reactant level, not from the height of the peak.

- Exothermic: products have lower potential energy and are more stable in this energetic sense than reactants; $\Delta H < 0$.
- Endothermic: products have higher potential energy and are less stable in this energetic sense; $\Delta H > 0$.
- The peak represents a barrier. Its height above the reactants relates to **activation energy**; it does not equal the enthalpy change. An exothermic fuel may still need ignition.

Bond breaking requires energy; bond formation releases energy. The net balance determines the sign. Do not say that burning releases energy because bonds are broken. Quantitative bond-energy calculations belong to the next unit. Nor does a negative enthalpy change, by itself, tell us how fast a reaction occurs or fully determine whether it is thermodynamically favourable.

Worked comparison. On an arbitrary common energy scale, reactants are at 80 kJ, products at 30 kJ and the peak at 140 kJ for one stated reaction amount. The endpoint change is $\Delta H = 30 - 80 = -50$ kJ: exothermic. The forward barrier is $140 - 80 = 60$ kJ. Reversing the reaction gives +50 kJ, not -60 kJ.

Standard enthalpy changes

A **standard enthalpy change** refers to a specified reaction at constant pressure with substances in their standard states. Standard pressure is 100 kPa (1 bar). A temperature must also be stated; tables commonly use 298 K. Standard does not mean 0 °C, and 298 K is not part of the definition of standard pressure.

Specify physical states; for a solution, the school-level standard concentration is 1 mol dm⁻³. Include state symbols in thermochemical equations: producing water vapour rather than liquid water changes the enthalpy change.

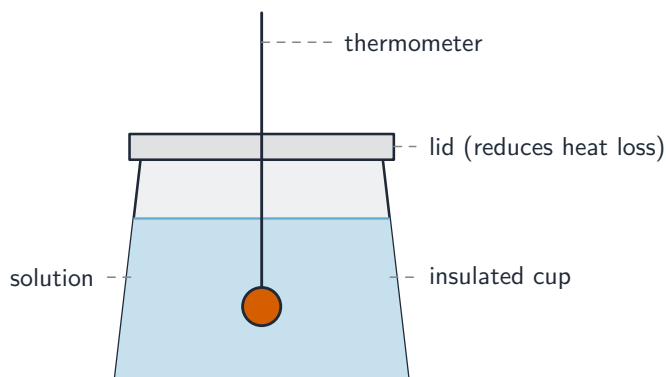
A quoted molar reaction enthalpy is tied to the equation as written. Double every coefficient and the heat change for that reaction amount doubles; reverse the equation and its sign reverses. Always explain the basis of “per mole”: per mole of fuel burned, salt dissolved, water formed, or reaction as written. A cup experiment at other conditions yields an experimental estimate; do not attach a standard symbol merely because a thermometer was used.

Calorimetry: from temperature to energy

Calorimetry measures heat transfer through a temperature change. For a substance whose specific heat capacity is approximately constant over the range,

$$q = mc\Delta T, \quad \Delta T = T_{\text{final}} - T_{\text{initial}}.$$

Here m is the mass being heated or cooled, c its specific heat capacity and q the heat it gains. Specific heat capacity is the energy needed to raise the temperature of unit mass by 1 K. With m in grams and $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$, q is in joules. A temperature *difference* of 1 °C equals 1 K: no 273 conversion is needed for a difference.



*Measure the mass and temperature change of the material that absorbs the heat;
reduce exchange with the room.*

For dilute aqueous mixtures we often approximate density as 1.00 g cm^{-3} and specific heat capacity as that of water. Use the **total** solution mass, not just the acid volume or the mass of solid added. If dissolving a solid and the stated model includes it, add its mass to the water mass. These are assumptions to justify, not universal properties of all solutions.

If heat exchange with the room and the cup's heat capacity are neglected,

$$q_{\text{reaction}} = -q_{\text{solution}}, \quad \Delta H_{\text{r}} = -\frac{q_{\text{solution}}}{1000 n_{\text{reaction}}}.$$

Convert J to kJ before division. The smallest reactant amount/coefficient ratio gives n_{reaction} ; for 1:1 equations, use the **limiting reactant** amount.

■ Worked example: neutralization with unequal amounts

Mix 25.0 cm³ of 2.00 mol dm⁻³ HCl with 40.0 cm³ of 1.00 mol dm⁻³ NaOH. Both start at 21.0 °C and the mixture reaches 28.0 °C. Use the water approximations above and neglect the cup.

The equation is $\text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}$.

$$n_{\text{HCl}} = cV = 2.00 \times 0.0250 = 0.0500 \text{ mol.}$$

$$n_{\text{NaOH}} = cV = 1.00 \times 0.0400 = 0.0400 \text{ mol.}$$

NaOH limits this 1:1 reaction; 0.0400 mol of water forms. The heated solution mass is $(25.0 + 40.0) \times 1.00 = 65.0 \text{ g}$, and $\Delta T = 7.0 \text{ K}$.

$$q_{\text{solution}} = mc\Delta T = 65.0 \times 4.18 \times 7.0 = 1901.9 \text{ J.}$$

$$\Delta H_{\text{r}} = -\frac{q_{\text{solution}}}{1000n} = -\frac{1901.9}{1000 \times 0.0400} \approx -48 \text{ kJ mol}^{-1}.$$

The solution gains heat; the reaction loses heat. Dividing by 0.0500 mol would include unreacted acid. The result is an experimental estimate, not evidence that heat loss was absent.

■ Worked example: an endothermic dissolution

Dissolve 4.00 g of a salt of molar mass 80.0 g mol⁻¹ in 100.0 g water. The temperature falls from 22.0 to 18.5 °C. Include the salt in the solution mass and assume $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$.

$$n = \frac{m_{\text{salt}}}{M} = \frac{4.00}{80.0} = 0.0500 \text{ mol.}$$

$$q_{\text{solution}} = mc\Delta T = 104.0 \times 4.18 \times (-3.5) = -1521.52 \text{ J.}$$

$$\Delta H_{\text{solution}} = -\frac{q_{\text{solution}}}{1000n} = -\frac{-1521.52}{1000 \times 0.0500} \approx +30 \text{ kJ mol}^{-1}.$$

Keep the negative temperature change through the calculation. The solution loses heat while the dissolving process gains it.

■ If the cup absorbs appreciable heat

A calibrated calorimeter heat capacity C_{cal} is in J K^{-1} , so it needs no mass factor. When cup and solution undergo the same temperature change,

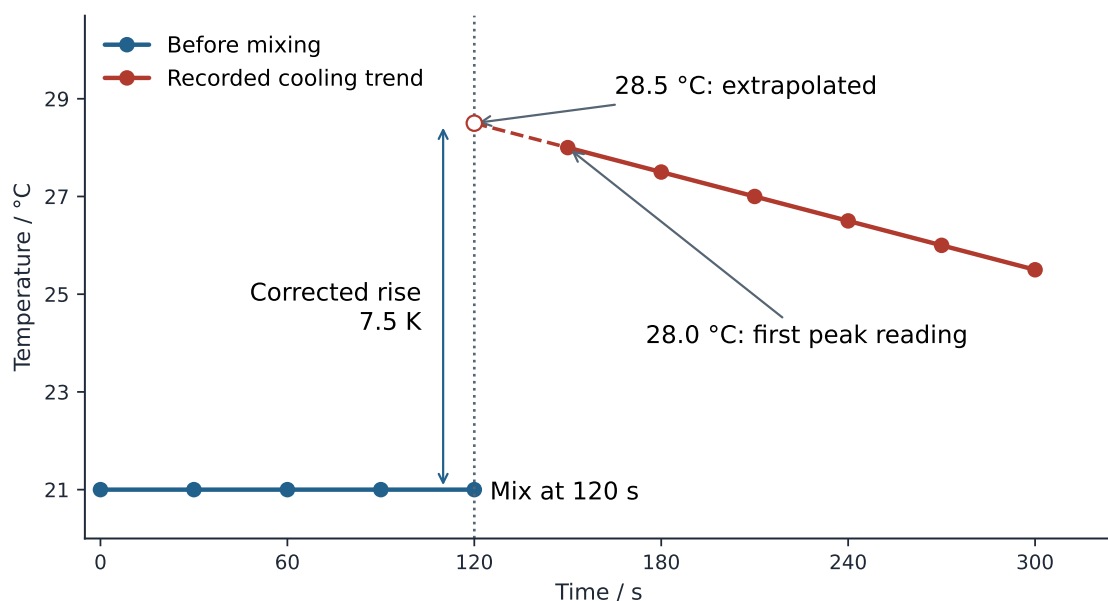
$$q_{\text{surroundings}} = (mc + C_{\text{cal}})\Delta T.$$

For 100.0 g solution, $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$, $C_{\text{cal}} = 32.0 \text{ J K}^{-1}$ and a 5.00 K rise, $q_{\text{surroundings}} = (418 + 32.0) \times 5.00 = 2250 \text{ J}$. If 0.0500 mol reacts, $\Delta H = -45.0 \text{ kJ mol}^{-1}$. Ignoring the cup gives $-41.8 \text{ kJ mol}^{-1}$, a smaller magnitude.

Designing and interpreting an experiment

Use approved apparatus, eye protection and teacher-specified reagents and quantities. An insulated cup and lid reduce heat exchange; a stirrer makes temperature more uniform. Immerse the thermometer bulb without touching the cup. Bring solutions to the same initial temperature, or account for their different starting temperatures; averaging them without regard to their masses and heat capacities is not generally valid.

Record temperature regularly before mixing, note the mixing time, then record through the rise or fall and subsequent return towards room temperature. A single observed maximum may miss heat already lost while mixing and reading.



Extrapolate suitable pre- and post-reaction trends to the same mixing time; do not read the extrapolated value at a later time.

In this simplified invented example, mixing is at 120 s, the baseline is 21.0 °C, and the later cooling trend extrapolates to 28.5 °C at mixing. The corrected rise is $28.5 - 21.0 = 7.5 \text{ K}$ rather than the observed $28.0 - 21.0 = 7.0 \text{ K}$. If all else is unchanged, the corrected heat and enthalpy magnitude are $7.5/7.0 \approx 1.07$ times larger. The cooling trend must represent heat exchange during the short reaction. Slow reaction, evaporation or changing conditions can invalidate that assumption.

For combustion, heat water above a spirit burner and weigh the capped burner before and after. The fuel amount comes from mass *lost* divided by molar mass; the heated

mass in $mc\Delta T$ is the water. Keep the burner–container distance consistent and reduce draughts without restricting oxygen. A flame shield must not turn the apparatus into a sealed vessel. Extinguish before weighing; never refill a hot burner.

Heat escaping to the air, heat absorbed by apparatus and incomplete combustion commonly make a measured combustion enthalpy less negative than the complete-combustion value. Soot suggests incomplete combustion. Fuel evaporation increases apparent fuel use without supplying the measured heat and also reduces the calculated magnitude. Repetition reduces random uncertainty, not systematic heat-loss bias.

Distinguish **uncertainty** from percentage disagreement with a reference. If each temperature reading has a stated maximum uncertainty of $\pm 0.1\text{ }^{\circ}\text{C}$, conservative addition gives $\pm 0.2\text{ K}$ in the difference. A 2.0 K rise has 10% temperature-difference uncertainty; a 10.0 K rise has 2%. Do not increase quantities beyond safe limits merely to improve this percentage.

Practice

Use $c = 4.18\text{ J g}^{-1}\text{ K}^{-1}$ unless stated otherwise. Neglect apparatus heat capacity except in question 8. All data are invented.

1. Explain the difference between heat and temperature. Compare the heat needed to warm 50 g and 100 g of water by the same 4 K .
2. A reaction cools its solution. State the signs of q_{solution} and $\Delta H_{\text{reaction}}$, and identify the direction of energy transfer.
3. An energy profile has reactants at 40 kJ , products at 90 kJ and a peak at 130 kJ for a stated reaction amount. Find the forward enthalpy change, forward barrier and reverse enthalpy change. State which endpoint is lower in energy.
4. A student calls any reaction measured at $25\text{ }^{\circ}\text{C}$ “standard” and omits state symbols. Explain two problems with this claim.
5. A 75.0 g solution warms from 20.0 to $24.0\text{ }^{\circ}\text{C}$. Calculate the heat it gains and the molar reaction enthalpy if 0.0250 mol reacts.
6. Mix 30.0 cm^3 of $1.00\text{ mol dm}^{-3}\text{ HCl}$ with 20.0 cm^3 of $1.00\text{ mol dm}^{-3}\text{ NaOH}$. Both start at $20.0\text{ }^{\circ}\text{C}$; the final temperature is $25.0\text{ }^{\circ}\text{C}$. Use density 1.00 g cm^{-3} . Determine the limiting reactant and enthalpy change per mole of water formed.
7. A 2.00 g salt, molar mass 100 g mol^{-1} , dissolves in 50.0 g water. Temperature falls by 2.0 K . Include the salt in the heated/cooled mass. Calculate the molar enthalpy of solution.
8. A 100.0 g solution and calorimeter of heat capacity 40.0 J K^{-1} warm by 4.00 K when 0.0400 mol reacts. Calculate the enthalpy change including the calorimeter. Explain the effect of ignoring it.
9. An alcohol of molar mass 46.0 g mol^{-1} loses 0.460 g from a burner while 200.0 g water warms by 12.0 K . Calculate the experimental molar combustion enthalpy. Explain why using 0.460 g in $mc\Delta T$ is wrong.

10. A reaction starts at 20.0 °C. The observed peak is 26.0 °C; extrapolation to mixing gives 26.8 °C. Find both temperature rises and the factor by which the calculated enthalpy magnitude changes.
11. Give two separate reasons why a spirit-burner result may be less negative than a complete-combustion reference, and one appropriate improvement for each. Would repeating the unmodified experiment remove these biases?
12. Two temperature readings each have a maximum uncertainty of ± 0.1 °C. Find the conservative percentage uncertainty in a 3.0 K rise. Distinguish this from percentage disagreement when an experimental enthalpy is $-42.0 \text{ kJ mol}^{-1}$ and the stated reference is $-50.0 \text{ kJ mol}^{-1}$.
13. In the equation $\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$, the stated enthalpy change is -286 kJ per mole of reaction. Find the heat transferred by the reaction when 0.250 mol water forms, then state the enthalpy changes for the doubled equation and the reversed original equation.

■ Worked answers

1. Heat is energy transferred because of a temperature difference; temperature relates to average particle kinetic energy. $q = mc\Delta T$: the 50 g sample needs $50 \times 4.18 \times 4 = 836 \text{ J}$ and the 100 g sample needs 1672 J. Doubling the mass doubles the required heat, although both temperature rises are equal.
2. $q_{\text{solution}} < 0$, so $q_{\text{reaction}} > 0$ and $\Delta H_{\text{reaction}} > 0$: endothermic. Energy transfers from the thermal surroundings to the reacting system.
3. $\Delta H = H_{\text{products}} - H_{\text{reactants}} = 90 - 40 = +50 \text{ kJ}$; forward barrier = $130 - 40 = 90 \text{ kJ}$. The reverse enthalpy change is -50 kJ . Reactants are lower in potential energy; the forward reaction is endothermic.
4. Temperature alone does not establish standard states: pressure, solution concentration where applicable and physical states matter. State symbols distinguish different initial and final states, whose enthalpies differ. The temperature must be specified too.
5. $q = mc\Delta T = 75.0 \times 4.18 \times 4.0 = 1254 \text{ J}$. $\Delta H = -q/(1000n) = -1254/(1000 \times 0.0250) \approx -50 \text{ kJ mol}^{-1}$; the reaction supplies the heat gained by solution.
6. $n = cV$: HCl is $1.00 \times 0.0300 = 0.0300 \text{ mol}$; NaOH is $1.00 \times 0.0200 = 0.0200 \text{ mol}$. NaOH limits the 1:1 reaction, giving 0.0200 mol water. Total mass is 50.0 g. $q = mc\Delta T = 50.0 \times 4.18 \times 5.0 = 1045 \text{ J}$. $\Delta H = -q/(1000n) = -1045/20.0 \approx -52 \text{ kJ mol}^{-1}$.
7. $n = m/M = 2.00/100 = 0.0200 \text{ mol}$. $q = mc\Delta T = 52.0 \times 4.18 \times (-2.0) = -434.72 \text{ J}$. $\Delta H = -q/(1000n) = +434.72/20.0 \approx +22 \text{ kJ mol}^{-1}$. The dissolution absorbs energy.
8. $q = (mc + C_{\text{cal}})\Delta T = (100.0 \times 4.18 + 40.0) \times 4.00 = 1832 \text{ J}$. $\Delta H = -q/(1000n) = -1832/40.0 = -45.8 \text{ kJ mol}^{-1}$. Ignoring the calorimeter gives $-41.8 \text{ kJ mol}^{-1}$, underestimating the magnitude because it omits energy gained by the cup.

9. $q = mc\Delta T = 200.0 \times 4.18 \times 12.0 = 10032 \text{ J}$. $n = m_{\text{fuel}}/M = 0.460/46.0 = 0.0100 \text{ mol}$. $\Delta H = -q/(1000n) = -10032/10.0 \approx -1.00 \times 10^3 \text{ kJ mol}^{-1}$. The water is the material whose temperature rise is measured; fuel mass instead determines how many moles burned.
10. $\Delta T_{\text{observed}} = 26.0 - 20.0 = 6.0 \text{ K}$; $\Delta T_{\text{corrected}} = 26.8 - 20.0 = 6.8 \text{ K}$. The magnitude increases by a factor $6.8/6.0 \approx 1.13$ (about 13%); the exothermic sign stays negative.
11. Heat escapes to the room: insulation or a suitable draught shield reduces that transfer. Incomplete combustion produces less heat than complete combustion: ensure an adequate oxygen supply and inspect for soot. Repetition alone does not eliminate either systematic bias; a shield that restricts oxygen could worsen combustion.
12. Conservative uncertainty in the difference is $0.1 + 0.1 = 0.2 \text{ K}$, so percentage uncertainty is $(0.2/3.0) \times 100 \approx 6.7\%$. Percentage disagreement is $|-42.0 - (-50.0)|/50.0 \times 100 = 16\%$. The first quantifies measurement uncertainty; the second compares a result with a specified reference and may reflect systematic effects.
13. $q_{\text{reaction}} = n\Delta H = 0.250 \times (-286) = -71.5 \text{ kJ}$: 71.5 kJ transfers to the surroundings. For $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O(l)}$, the enthalpy change is -572 kJ per mole of the doubled reaction. Reversing the original equation gives $+286 \text{ kJ}$ per mole of that reversed reaction.

Exam tips

- Balance the equation and find the limiting amount; use all the solution mass in $mc\Delta T$. Name the system before assigning a sign: warming solution means a negative reaction enthalpy in this model.
- Show the J-to-kJ conversion, retain guard digits and state the mole basis. A standard symbol needs standard conditions and states; a negative sign does not prove a fast reaction.
- Label profile axes and draw ΔH between endpoints. Explain an experimental error through its cause, direction of bias and a specific improvement; “human error” is insufficient.

Reference: IB Chemistry guide, Reactivity 1.1.