Measuring enthalpy changes

Reactivity 1.1 · shared SL and HL content

Why does the cup warm?

Track energy, then calculate an enthalpy change

Two solutions react and their temperature rises. Choose the chemical reaction as the system: where has the energy gone?

This unit links energy conservation, signed profiles, standard states and calorimetry. Reactivity 1.1 is shared by SL and HL.

Use balanced equations, limiting reactants and $n = cV$ or $n = m/M$. All experimental numbers in this deck are invented.

Heat and temperature answer different questions

Temperature

Relates to average particle kinetic energy.

50 g and 100 g water may have the same temperature.

Heat transfer

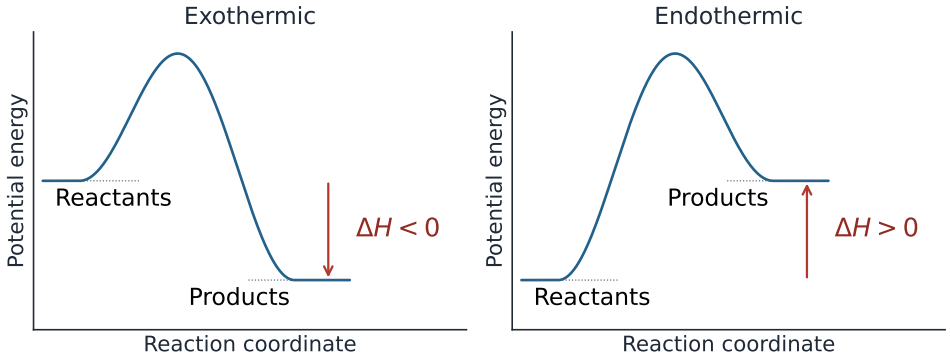
Energy transferred because of a temperature difference, measured in J.

For the same rise, 100 g water needs twice the heat of 50 g.

Whose sign is it?

Reaction	**Solution**	**Reaction enthalpy**
Exothermic	Warms: gains heat	$\Delta H < 0$
Endothermic	Cools: loses heat	$\Delta H > 0$

Read the endpoints, not the peak



> Potential energy against reaction coordinate: products minus reactants gives the sign.

A peak is not an enthalpy change

1. **Given.** Reactants 80 kJ; products 30 kJ; peak 140 kJ, for one reaction amount.
2. **Endpoints.** $\Delta H = H_{\text{products}} - H_{\text{reactants}} = 30 - 80 = -50$ kJ.
3. **Barrier.** Forward barrier $= 140 - 80 = 60$ kJ; reverse enthalpy change $= +50$ kJ.

Decision

Lower-energy products do not imply a fast reaction. A fuel can still need ignition.

Check the direction

Draw before looking at the answer

Reactants are at 40 kJ, products at 90 kJ and the peak at 130 kJ. Sketch and label the profile.

Find the forward enthalpy change, forward barrier and reverse enthalpy change. Which endpoint is lower in energy?

Profile reasoning

1. **Endpoint difference.** $\Delta H = 90 - 40 = +50$ kJ: **endothermic**.
2. **Barrier.** $E_{\mathrm{a,forward}} = 130 - 40 = 90$ kJ.
3. **Reverse.** $\Delta H_{\mathrm{reverse}} = -50$ kJ; the original reactants are lower in potential energy.

Decision

Bond breaking absorbs energy; bond formation releases it. The net balance sets the sign.

What does standard specify?

Temperature alone is insufficient

Use standard states at 100 kPa (1 bar), with temperature specified; tables commonly use 298 K.

Give physical states. At school level, standard solution concentration is 1 mol dm^{-3}. Water liquid and water vapour are different states.

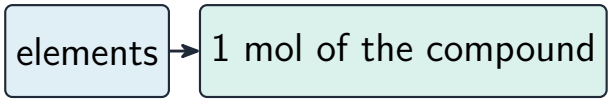
A measured cup result is an experimental estimate. A thermometer reading of 25 °C alone does not justify a standard symbol.

The equation fixes the amount basis

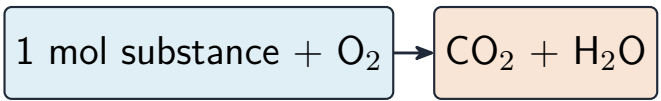
1. **Given.** $\mathrm{H_2(g)} + \frac{1}{2}\mathrm{O_2(g)} \rightarrow \mathrm{H_2O(l)}$: $\Delta H = -286$ kJ per mole of reaction.
2. **A smaller amount.** $q = n\Delta H = 0.250 \times (-286) = -71.5$ kJ when 0.250 mol water forms.
3. **Change the equation.** Double coefficients: -572 kJ per mole of doubled reaction. Reverse original: $+286$ kJ per mole of reversed reaction.

Formation and combustion

formation ΔH_f:

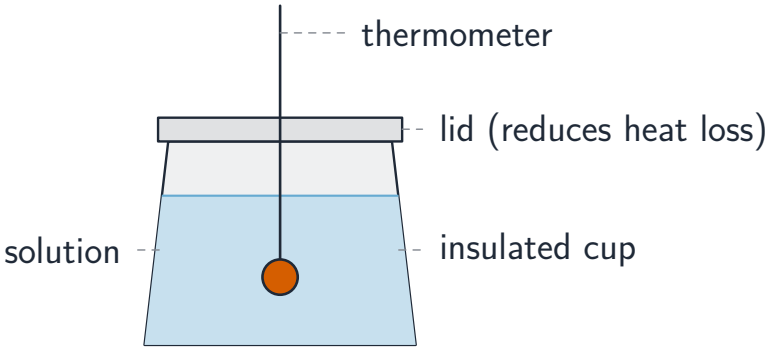


combustion ΔH_c:



Both under standard conditions

Measure the material that changes temperature



A lid and insulation reduce heat exchange; immerse the bulb in the solution.

From thermometer readings to joules

1. **Model.** $q = mc\Delta T$; $\Delta T = T_{\text{final}} - T_{\text{initial}}$.
2. **Units.** Use solution mass in g, $c = 4.18$ J g^{-1} K^{-1} and a temperature difference in K.
3. **Boundary.** $q_{\text{reaction}} = -q_{\text{solution}}$ when other heat transfers are neglected.

Decision

Use the total solution mass. A 1 °C temperature difference equals 1 K.

Neutralization: decide the reacting amount first

Invented experiment

Mix 25.0 cm^3 of 2.00 mol dm^{-3} HCl with 40.0 cm^3 of 1.00 mol dm^{-3} NaOH. Both start at 21.0 °C; final temperature 28.0 °C.

Use density 1.00 g cm^{-3} and $c = 4.18$ J g^{-1} K^{-1}; neglect the cup. Find limiting reactant, heated mass and enthalpy per mole of water.

$$\mathrm{HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)}$$

The excess acid does not belong in the denominator

1. **HCl amount.** $n = cV = 2.00 \times 0.0250 = 0.0500$ mol.
2. **NaOH amount.** $n = cV = 1.00 \times 0.0400 = 0.0400$ mol.
3. **Reacting amount.** The equation is 1:1. NaOH is the **limiting reactant**; 0.0400 mol water forms.

Decision

For other equations, compare available amount divided by stoichiometric coefficient.

Neutralization: heat and sign

1. **Heated solution.** $m = (25.0 + 40.0) \times 1.00 = 65.0$ g; $\Delta T = 28.0 - 21.0 = 7.0$ K.
2. **Heat gained.** $q = mc\Delta T = 65.0 \times 4.18 \times 7.0 = 1901.9$ J.
3. **Molar enthalpy.** $\Delta H = -q/(1000n) = -1901.9/(1000 \times 0.0400) \approx -48$ kJ mol^{-1}.

Decision

The solution gains heat; the reacting system loses it.

Cooling means a positive reaction enthalpy

1. **Invented data.** 4.00 g salt, molar mass 80.0 g mol^{-1}, dissolves in 100.0 g water; 22.0 $\rightarrow$ 18.5 °C.
2. **Amount and heat.** $n = 4.00/80.0 = 0.0500$ mol; $q = 104.0 \times 4.18 \times (-3.5) = -1521.52$ J.
3. **Sign.** $\Delta H = -q/(1000n) = +1521.52/50.0 \approx +30$ kJ mol^{-1}.

Decision

Include the salt in the solution mass under this stated model.

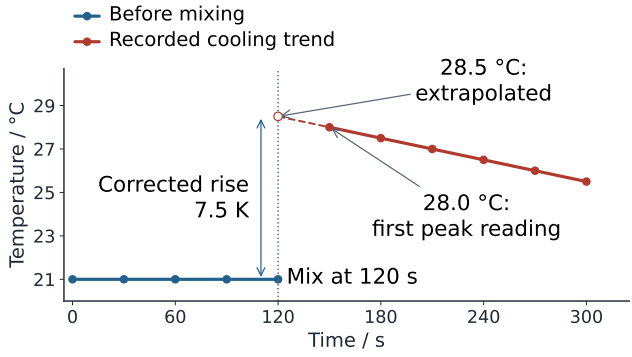
Account for the calorimeter

1. **Heat capacities.** 100.0 g solution: $mc = 418$ J K^{-1}. Cup: $C_{\mathrm{cal}} = 32.0$ J K^{-1}.
2. **Heat gained.** $q = (mc + C_{\mathrm{cal}})\Delta T = (418 + 32.0) \times 5.00 = 2250$ J.
3. **Reaction amount.** For 0.0500 mol, $\Delta H = -2250/50.0 = -45.0$ kJ mol^{-1}.

Decision

Omitting the cup gives -41.8 kJ mol^{-1}, underestimating the magnitude.

Extrapolate to the mixing time



Invented data: corrected rise $= 28.5 - 21.0 = 7.5$ K, rather than the observed 7.0 K.

What the correction assumes

An extrapolation is still a model

Record a baseline, note mixing time and follow the post-reaction trend. Compare both trends at that same mixing time.

The later cooling trend must reasonably represent heat loss during the short reaction. Slow reaction, evaporation or changing conditions can spoil that estimate.

Use an insulated cup and lid, stir, and start solutions at a common temperature. Do not average unequal starting temperatures without considering heat capacities.

Combustion: two different masses

Separate heat gained from fuel consumed

Weigh the capped spirit burner before and after: mass lost divided by fuel molar mass gives moles burned.

In $q = mc\Delta T$, use the water mass because its temperature rise is measured. Fuel mass is not the heated water mass.

Use approved quantities and adequate ventilation. Extinguish before weighing; never refill a hot burner or heat a sealed vessel.

Try a fuel calculation

Invented data: work before revealing

Burner mass decreases by 0.460 g. Fuel molar mass is 46.0 g mol^{-1}. 200.0 g water warms by 12.0 K; use $c = 4.18$ J g^{-1} K^{-1}.

Find heat gained by water, fuel amount and experimental molar combustion enthalpy. Then explain the direction of heat-loss bias.

Fuel calculation and its limit

1. **Heat.** $q = mc\Delta T = 200.0 \times 4.18 \times 12.0 = 10032$ J.
2. **Fuel.** $n = m/M = 0.460/46.0 = 0.0100$ mol.
3. **Enthalpy.** $\Delta H = -q/(1000n) = -10032/10.0 \approx -1.00 \times 10^{3}$ kJ mol^{-1}.

Decision

Escaping heat makes the estimate less negative: the water receives only part of the released energy.

Match each limitation to its effect

Limitation	Effect on magnitude	Improvement
Heat escapes	Too small	Insulation; suitable shield
Incomplete burning	Too small	Adequate oxygen; inspect soot
Fuel evaporates	Too small	Cap; weigh after extinguishing

Uncertainty is not disagreement

1. **Reading uncertainty.** Each reading ± 0.1 °C gives conservative difference **uncertainty** ± 0.2 K.
2. **A 3.0 K rise.** Percentage uncertainty $= (0.2/3.0) \times 100 \approx 6.7\%$.
3. **A reference comparison.** -42.0 versus -50.0 kJ mol^{-1}: disagreement $= (8.0/50.0) \times 100 = 16\%$.

Decision

The 16% disagreement may include systematic bias; it is not the thermometer uncertainty.

Independent transfer

Solve, then diagnose

Mix 30.0 cm^3 1.00 mol dm^{-3} HCl with 20.0 cm^3 1.00 mol dm^{-3} NaOH, both at 20.0 °C. Final temperature: 25.0 °C.

Use the 1:1 equation, density 1.00 g cm^{-3} and $c = 4.18$ J g^{-1} K^{-1}. Neglect cup heat capacity. Find ΔH per mole water.

Explain the effect of using only the acid mass, then try questions 1–13 in the handout.

Independent transfer: check the reasoning

1. **Amounts.** HCl 0.0300 mol; NaOH 0.0200 mol: NaOH limits, so 0.0200 mol water forms.
2. **Heat gained.** $q = mc\Delta T = 50.0 \times 4.18 \times 5.0 = 1045$ J.
3. **Enthalpy.** $\Delta H = -q/(1000n) = -1045/20.0 \approx -52$ kJ mol^{-1}.

Decision

Using only 30.0 g underestimates heat gained and the enthalpy magnitude.

Check yourself

Five checks

System and sign: who gains heat? Equation and amount: which reactant limits?

Mass: all the heated solution? Units: J converted to kJ? Final answer: kJ per stated mole basis?

For evaluation, name a cause, its direction of bias and a feasible improvement. For profiles, label both axes and the endpoint difference.