



AQA CHEMISTRY 7405

ENERGETICS I

ENTHALPY CHANGES AND CALORIMETRY

HEAT TRANSFER • CALCULATIONS • REQUIRED PRACTICAL 2

			
SPEC	CRACK IT	TEST IT	DONE WHEN

From a temperature reading to an enthalpy change

Learn how to interpret heat transfer, define standard enthalpy changes and calculate energy per mole. Then measure and evaluate enthalpy changes using solution and combustion calorimetry.

AQA SECTION	THIS PACK TEACHES
3.1.4.1	Exothermic/endothermic changes; standard conditions; standard formation and combustion enthalpies.
3.1.4.2 and RP2	$q = mc\Delta T$; molar enthalpy calculations; practical methods, temperature-time correction and evaluation.
Short supporting link	Enthalpy profiles and activation energy (3.1.5.1); mole ratios and practical measurement skills.

BRING FORWARD $n = m \div M$; $n = cV$ with V in dm^3 ; balanced-equation mole ratios. Here c in $q = mc\Delta T$ means specific heat capacity, not concentration.

HOW TO USE Read each example before attempting its different TEST IT question. Show equations, substitutions and units. Use the labelled graph/drawing space and full-width dotted lines. Worked answers are at the back.

Energetics II follows with Hess's law and mean bond enthalpies. This pack focuses on direct enthalpy measurements and the definitions needed to interpret them.

Keep unrounded working values. Use the precision requested in each question; otherwise choose a sensible number of significant figures from the supplied data. Molar masses and heat capacities are supplied where needed.

Specification checked 14 October 2026. Original questions and worked examples, with complete answers.

1. Track the heat transfer



SPEC 3.1.4.1 Explain exothermic and endothermic changes and the sign of ΔH .



CRACK IT The sign belongs to the reacting system. Use the temperature of the surroundings to work out which way heat moved.

FULL-MARK DEFINITION Enthalpy change, ΔH , is the heat energy change at constant pressure. The system is the reaction being studied; the surroundings are everything exchanging energy with it.

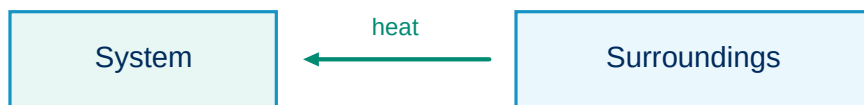
EXOTHERMIC

$$\Delta H < 0$$



ENDOTHERMIC

$$\Delta H > 0$$



CHANGE	HEAT TRANSFER	SURROUNDINGS	SIGN OF ΔH
Exothermic	System $\rightarrow$ surroundings	Warm up	Negative
Endothermic	Surroundings $\rightarrow$ system	Cool down	Positive

Worked examples: follow the energy

A neutralisation warms its solution from 20.0 °C to 27.0 °C. The solution gains heat from the reacting system, so the reaction is exothermic and ΔH is negative.

Dissolving ammonium nitrate cools the solution. Energy is transferred from the surroundings into the dissolving system, so the change is endothermic and ΔH is positive.

EXAM TRAP A temperature rise does not mean a positive reaction enthalpy. A warmer solution has gained heat that the reaction has released.



TEST IT 1

a) A displacement reaction warms its solution from 19.5 °C to 25.5 °C. State the type of change, the direction of heat transfer and the sign of ΔH . [3 marks]

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b) A different reaction cools the surroundings. Explain whether the enthalpy of the products is higher or lower than that of the reactants. [2 marks]

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✓ **DONE WHEN** I can link temperature, heat flow and the sign of ΔH .

2. Read and draw enthalpy profiles

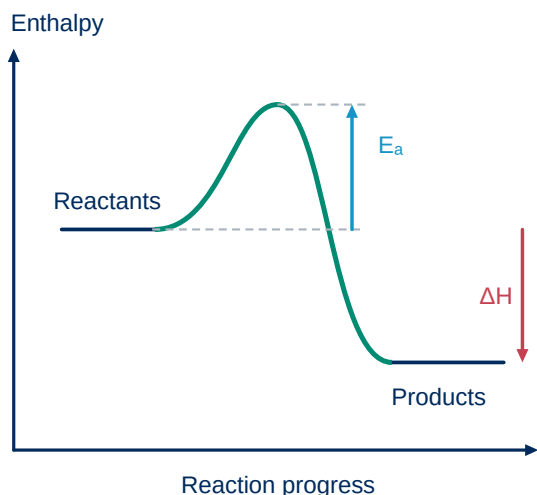


SPEC 3.1.4.1; link 3.1.5.1 Show reactants, products, ΔH and the activation energy on an enthalpy profile.

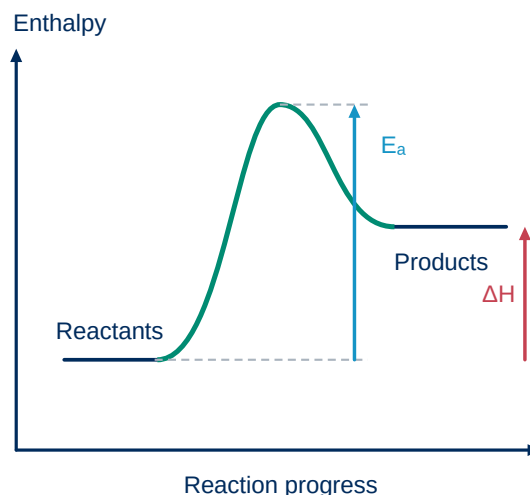


CRACK IT ΔH compares the two levels. Activation energy compares the reactant level with the top of the barrier.

KEY TERMS $\Delta H = H(\text{products}) - H(\text{reactants})$. Activation energy, E_a , is the minimum energy particles need for a successful collision. This is a short link to AQA 3.1.5.1.



Exothermic: ΔH negative



Endothermic: ΔH positive

Label the vertical axis “Enthalpy” and the horizontal axis “Reaction progress”. The horizontal axis is not time. Draw the ΔH arrow from the reactant level to the product level. E_a runs upward from the reactant level to the peak.

Worked example: read numerical levels

Reactants are at 80 kJ mol^{-1} , products at 30 kJ mol^{-1} and the peak at 140 kJ mol^{-1} , on the same relative scale. $\Delta H = 30 - 80 = -50 \text{ kJ mol}^{-1}$; forward $E_a = 140 - 80 = 60 \text{ kJ mol}^{-1}$. The reaction is exothermic even though it needs an energy input to start.



TEST IT 2

a) Draw an enthalpy profile with reactants at 20 kJ mol^{-1} , products at 65 kJ mol^{-1} and a peak at 100 kJ mol^{-1} . Label both axes, the reactants, products, ΔH and forward E_a . State the two energy values. [4 marks]

2(a) Draw and label the endothermic profile and energy values.

b) Explain why an exothermic reaction can still require heating to begin. [1 mark]

✓ **DONE WHEN** I can distinguish the overall enthalpy change from the activation energy.

3. Use standard conditions and state symbols



SPEC 3.1.4.1 Interpret standard enthalpy notation and specify the conditions for comparisons.



CRACK IT Standard means 100 kPa and a stated temperature. Always include the physical states in the equation.

STANDARD CONDITIONS Use a pressure of 100 kPa and a stated temperature, usually 298 K. Substances are in their standard states: their reference physical forms under those conditions. For a standard solution, concentration is 1 mol dm⁻³.

SYMBOL / TERM	MEANING
ΔH	Enthalpy change
$\Delta_r H^\ominus$	Standard enthalpy change of reaction for the equation as written
$\Delta_f H^\ominus$	Standard enthalpy change of formation
$\Delta_c H^\ominus$	Standard enthalpy change of combustion
kJ mol ⁻¹	Energy change per mole of the specified substance or reaction amount

Worked examples: the state matters



At 298 K, liquid water is the standard state of water. Changing the product to H₂O(g) describes a different energy change. Graphite, not diamond, is the reference form of carbon at these conditions.

For 2H₂(g) + O₂(g) → 2H₂O(l), a quoted reaction enthalpy refers to two moles of water formed per equation amount. It is not automatically an enthalpy of formation, whose basis is one mole of compound.

EXAM TRAP 298 K is commonly used, but “standard” does not mean the temperature can never be different. State the temperature given in the question.



TEST IT 3

a) State the standard pressure and usual stated temperature, and give the standard concentration of an aqueous solution. [3 marks]

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b) Two reactions convert ethanol into identical products under the same conditions. One starts with C₂H₅OH(l); the other starts with C₂H₅OH(g). Explain why their enthalpy changes differ. [2 marks]

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✓ **DONE WHEN** I can state the conditions, use state symbols and identify the mole basis.

4. Define standard enthalpy of formation



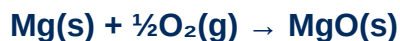
SPEC 3.1.4.1 Define $\Delta_f H^\ominus$ and write an equation for forming exactly one mole of a compound.



CRACK IT One mole of compound on the right. Only its elements, in their standard states, on the left.

FULL-MARK DEFINITION The standard enthalpy of formation is the enthalpy change when one mole of a compound is formed from its constituent elements, with all substances in their standard states under standard conditions.

Worked example 1: magnesium oxide



1. The product must be 1 mol of MgO, so its coefficient is 1.
2. Magnesium is a solid; oxygen is diatomic $\text{O}_2\text{(g)}$.
3. One oxygen atom is required per MgO, so use $\frac{1}{2}\text{O}_2$. Fractional coefficients are allowed.

Worked example 2: water



The formation equation follows the definition, even if the compound cannot be prepared directly in this single step. Formation can be exothermic or endothermic. The standard enthalpy of formation of an element in its standard state is zero by convention.

CHECK THE REACTANTS $\text{CO(g)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)}$ is not a formation equation: CO is a compound. Formation of CO_2 must start from C(s, graphite) and $\text{O}_2\text{(g)}$.



TEST IT 4

a) Define the standard enthalpy of formation. [3 marks]

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b) Write the standard formation equation for liquid ethanol, $\text{C}_2\text{H}_5\text{OH(l)}$, including state symbols. [2 marks]

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c) State $\Delta_f H^\ominus$ for $\text{O}_2\text{(g)}$ in its standard state. [1 mark]

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✓ **DONE WHEN** I can write a one-mole formation equation using elements in their standard states.

5. Define standard enthalpy of combustion



SPEC 3.1.4.1 Define $\Delta_c H^\ominus$ and write a complete-combustion equation for one mole of fuel.



CRACK IT One mole of the substance burns completely in oxygen. Oxygen does not have to have coefficient 1.

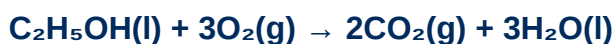
FULL-MARK DEFINITION The standard enthalpy of combustion is the enthalpy change when one mole of a substance burns completely in oxygen, with all reactants and products in their standard states under standard conditions.

Worked example 1: methane



Keep the methane coefficient at 1. Balance C, then H, then O. At 298 K the water product is liquid for this standard enthalpy definition.

Worked example 2: ethanol



There are seven oxygen atoms in the products. One is already in ethanol, so six more arrive as 3O_2 . The enthalpy is per mole of ethanol, not per mole of O_2 or CO_2 .

FORMATION	COMBUSTION
Makes 1 mol of a compound	Burns 1 mol of a substance completely
Starts from constituent elements	Uses oxygen as a reactant
Can be exothermic or endothermic	The fuel-combustion examples here are exothermic

EXAM TRAP Incomplete burning to CO or soot does not measure the enthalpy of complete combustion.



TEST IT 5

a) Define the standard enthalpy of combustion. [3 marks]

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b) Write the standard combustion equation for butane, $\text{C}_4\text{H}_{10}(\text{g})$, including state symbols. [2 marks]

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c) Why must the butane coefficient remain 1 for this definition? [1 mark]

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✓ **DONE WHEN** I can distinguish formation from combustion and preserve the correct one-mole basis.

6. Calculate the heat gained or lost



SPEC 3.1.4.2 Use $q = mc\Delta T$ with compatible units and a clearly identified heated material.



CRACK IT m is the mass whose temperature changes. It is usually the whole solution or the water, not the mass of the reacting solid or fuel.

$$q = mc\Delta T \quad \Delta T = T(\text{final}) - T(\text{initial})$$

SYMBOL	QUANTITY	UNIT USED HERE
q	Heat change of the measured water or solution	J
m	Mass of that water or solution	g
c	Specific heat capacity	$\text{J g}^{-1} \text{K}^{-1}$
ΔT	Temperature change	K or $^{\circ}\text{C}$

Specific heat capacity is the energy needed to raise the temperature of 1 g of a substance by 1 K. A temperature change of 1 $^{\circ}\text{C}$ equals a change of 1 K. AQA supplies c ; it does not need to be memorised.

Worked example 1: temperature rise

75.0 g of water warms from 21.20 $^{\circ}\text{C}$ to 29.60 $^{\circ}\text{C}$. Use $c = 4.18 \text{ J g}^{-1} \text{K}^{-1}$. $\Delta T = 8.40 \text{ K}$; $q = 75.0 \times 4.18 \times 8.40 = +2633.4 \text{ J} = +2.63 \text{ kJ}$. The positive sign means the water gains heat.

Worked example 2: temperature fall

60.0 g of solution cools by 3.20 K. $q(\text{solution}) = 60.0 \times 4.18 \times (-3.20) = -803 \text{ J}$. The solution loses heat, so the process causing the cooling absorbs it.

MASS FROM VOLUME If density is assumed to be 1.00 g cm^{-3} , $25.0 \text{ cm}^3 + 25.0 \text{ cm}^3$ gives 50.0 g of solution. Do not assume this density unless it is supplied or justified.



TEST IT 6

a) 90.0 g of water warms from 19.40 $^{\circ}\text{C}$ to 26.90 $^{\circ}\text{C}$. Calculate its heat change in J and kJ. Use $c = 4.18 \text{ J g}^{-1} \text{K}^{-1}$. [3 marks]

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b) 30.0 cm^3 and 20.0 cm^3 of aqueous solutions are mixed. Their density is 1.00 g cm^{-3} . State the mass used in $q = mc\Delta T$ and explain your choice. [2 marks]

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✓ **DONE WHEN** I can identify m , calculate ΔT and convert J to kJ.

7. Turn a heat change into a molar enthalpy



SPEC 3.1.4.2 Link q , moles and ΔH without confusing the solution with the reacting system.



CRACK IT Calculate the heat change of the surroundings, reverse its sign, then divide by the amount that actually reacted.

$$q(\text{reaction}) \approx -q(\text{surroundings}) \quad \Delta H = q(\text{reaction}) \div n$$

The approximation assumes negligible heat transfer to the apparatus and wider surroundings. With $q(\text{surroundings})$ in J, use $\Delta H = -q(\text{surroundings}) \div (1000 \times n)$ to obtain kJ mol^{-1} . State what the “per mole” refers to.

Worked example: dissolving a solid

2.00 g of NaOH ($M = 40.0 \text{ g mol}^{-1}$) dissolves in water. The final solution has mass 100.0 g and warms by 5.30 K. Take $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$. Find ΔH per mole of NaOH.

STAGE	SUBSTITUTION	RESULT
1 Heat of solution	$q = 100.0 \times 4.18 \times 5.30$	+2215.4 J
↓ 2 Amount dissolved	$n = 2.00 \div 40.0$	0.0500 mol
↓ 3 Sign and units	$q(\text{reaction}) = -2215.4 \div 1000$	-2.2154 kJ
↓ 4 Per mole	$\Delta H = -2.2154 \div 0.0500$	-44.3 kJ mol^{-1}

TWO DIFFERENT MASSES 100.0 g belongs in $q = mc\Delta T$ because the whole solution warms. 2.00 g belongs in $n = m \div M$ because that is the mass of NaOH dissolved.



TEST IT 7

a) 3.00 g of KNO_3 ($M = 101.1 \text{ g mol}^{-1}$) dissolves completely. The final solution has mass 100.0 g and cools by 2.50 K. Calculate ΔH per mole of KNO_3 . Use $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$ and give three significant figures. [5 marks]

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b) Explain the positive sign in your answer. [1 mark]

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✓ **DONE WHEN** I can keep the two masses separate and report a signed value in kJ mol^{-1} .

8. Use the mole ratio in neutralisation



SPEC 3.1.4.2 Calculate a neutralisation enthalpy using the limiting reagent and moles of water formed.



CRACK IT For a neutralisation value per mole of water, divide by $n(\text{H}_2\text{O})$, not automatically by $n(\text{acid})$.

NEUTRALISATION BASIS Enthalpy change of neutralisation is the enthalpy change when an acid and a base react to form one mole of water. This is a useful calorimetry application. $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$.

Worked example: a 1:2 acid-to-base ratio



Mix 25.0 cm^3 of $1.00 \text{ mol dm}^{-3} \text{ H}_2\text{SO}_4$ with 40.0 cm^3 of $1.00 \text{ mol dm}^{-3} \text{ NaOH}$, initially at the same temperature. The rise is 8.30 K . Use density = 1.00 g cm^{-3} and $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$. Assume complete neutralisation of the available hydroxide.

QUANTITY	H_2SO_4	NaOH	H_2O
Coefficient	1	2	2
1 $n = cV/1000$	0.0250 mol	0.0400 mol	Find from ratio
2 $n \div \text{coefficient}$	0.0250	0.0200 $\rightarrow$ limiting	—
3 Amount reacting / formed	0.0200 mol	0.0400 mol	$0.0400 \times 2/2 = 0.0400 \text{ mol}$

4. $m(\text{solution}) = 25.0 + 40.0 = 65.0 \text{ g}$. $q(\text{solution}) = 65.0 \times 4.18 \times 8.30 = 2255.11 \text{ J}$.

5. ΔH per mole of $\text{H}_2\text{O} = -2.25511 \div 0.0400 = -56.4 \text{ kJ mol}^{-1}$. Using moles of H_2SO_4 reacted would give the value per mole of acid, which is twice as large.

RATIO CHECK Compare $n \div \text{coefficient}$ for each reactant. The smaller value identifies the limiting reactant. Then use the equation to find the required mole basis.



TEST IT 8

a) 40.0 cm^3 of $0.500 \text{ mol dm}^{-3} \text{ HNO}_3$ is mixed with 30.0 cm^3 of $0.500 \text{ mol dm}^{-3} \text{ KOH}$. The temperature rises by 2.90 K . Calculate ΔH per mole of water formed. Use density = 1.00 g cm^{-3} , $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$ and a 1:1 neutralisation ratio. [5 marks]

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✓ **DONE WHEN** I can find the limiting reactant and use moles of water for neutralisation.

9. Apply calorimetry to solids



SPEC 3.1.4.2 Calculate enthalpy changes for displacement and dissolution using the stated assumptions.



CRACK IT First identify what reacts and what changes temperature. The solid mass may determine n while the solution mass determines q .

Worked example 1: displacement



Excess Zn is added to 50.0 cm³ of 0.500 mol dm⁻³ CuSO₄. The rise is 22.0 K. Use 50.0 g as the heated mass and $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$; neglect the solid and apparatus heat capacities.

1. CuSO₄ is limiting: $n = 0.500 \times 50.0 \div 1000 = 0.0250 \text{ mol}$.
2. $q(\text{solution}) = 50.0 \times 4.18 \times 22.0 = 4598 \text{ J}$.
3. ΔH per mole of CuSO₄ reacted = $-4.598 \div 0.0250 = -184 \text{ kJ mol}^{-1}$.

Worked example 2: include the dissolved solid in the mass

4.00 g of NH₄Cl ($M = 53.5 \text{ g mol}^{-1}$) dissolves in 100.0 g of water and cools it by 2.70 K. Take the final solution mass as 104.0 g and $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$.

$q(\text{solution}) = 104.0 \times 4.18 \times (-2.70) = -1173.744 \text{ J}$; $n(\text{NH}_4\text{Cl}) = 4.00 \div 53.5 = 0.0747664 \text{ mol}$. $\Delta H = +1.173744 \div 0.0747664 = +15.7 \text{ kJ mol}^{-1}$.

FOLLOW THE MODEL Use the heated mass and heat-capacity assumptions stated in the question. Do not automatically add every solid mass: an undissolved metal does not have the same c as water.



TEST IT 9

a) Excess Mg reacts with 40.0 cm³ of 0.250 mol dm⁻³ CuSO₄ by $\text{Mg} + \text{CuSO}_4 \rightarrow \text{MgSO}_4 + \text{Cu}$. The rise is 25.0 K. Use 40.0 g as the heated mass and $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$. Calculate ΔH per mole of CuSO₄ reacted. [4 marks]

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b) A soluble solid of mass 5.00 g is completely dissolved in 80.0 g of water. State the final solution mass used if its heat capacity is modelled using the supplied solution value of c . [1 mark]

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✓ **DONE WHEN** I can choose the reacting amount and the heated mass for each experiment.

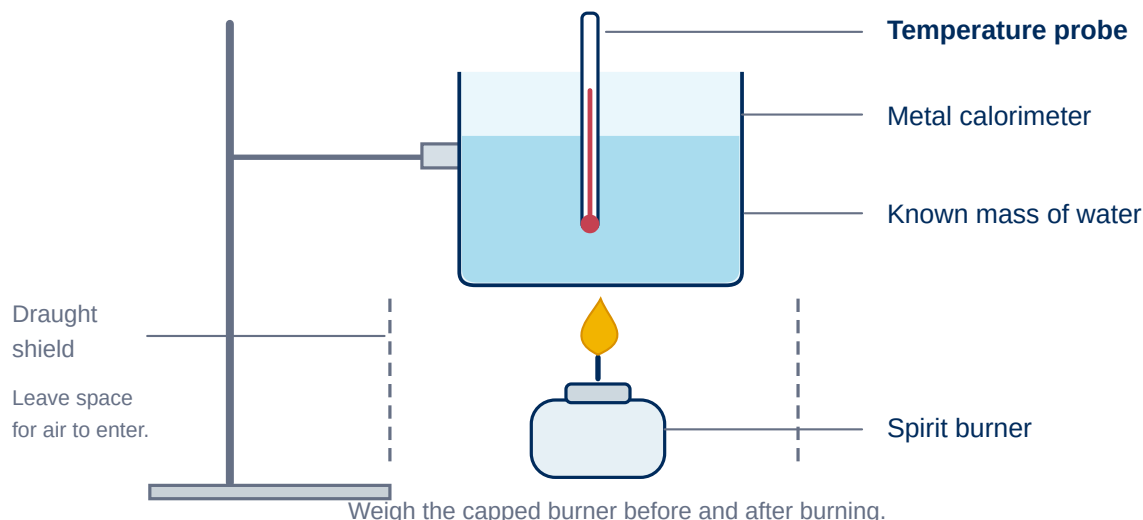
10. Measure a combustion enthalpy



SPEC 3.1.4.2 Use the mass of fuel burned and the heat gained by water to estimate $\Delta_c H$.



CRACK IT Water mass gives q . Fuel mass lost gives n . Keep these two measurements separate.



Worked example: ethanol heats water

A burner loses 0.600 g of ethanol ($M = 46.0 \text{ g mol}^{-1}$). 100.0 g of water warms by 35.0 K. Use $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$.

- $q(\text{water}) = 100.0 \times 4.18 \times 35.0 = 14\,630 \text{ J} = 14.630 \text{ kJ}$.
- $n(\text{ethanol}) = 0.600 \div 46.0 = 0.0130435 \text{ mol}$.
- Estimated $\Delta_c H = -14.630 \div 0.0130435 = -1.12 \times 10^3 \text{ kJ mol}^{-1}$.

WHAT THIS MEASURES An open-flame experiment estimates combustion enthalpy. It usually misses some released heat and does not reproduce every standard-state condition. Do not label the result as an accurate standard value without evidence.



TEST IT 10

a) A propan-1-ol burner decreases from 142.36 g to 141.61 g. It heats 150.0 g of water from 20.00 °C to 44.00 °C. Calculate the experimental combustion enthalpy per mole of fuel. Use $M = 60.0 \text{ g mol}^{-1}$ and $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$. [5 marks]

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b) Why is the water mass used in $q = mc\Delta T$? [1 mark]

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✓ **DONE WHEN** I can use mass loss for fuel moles and water measurements for heat.

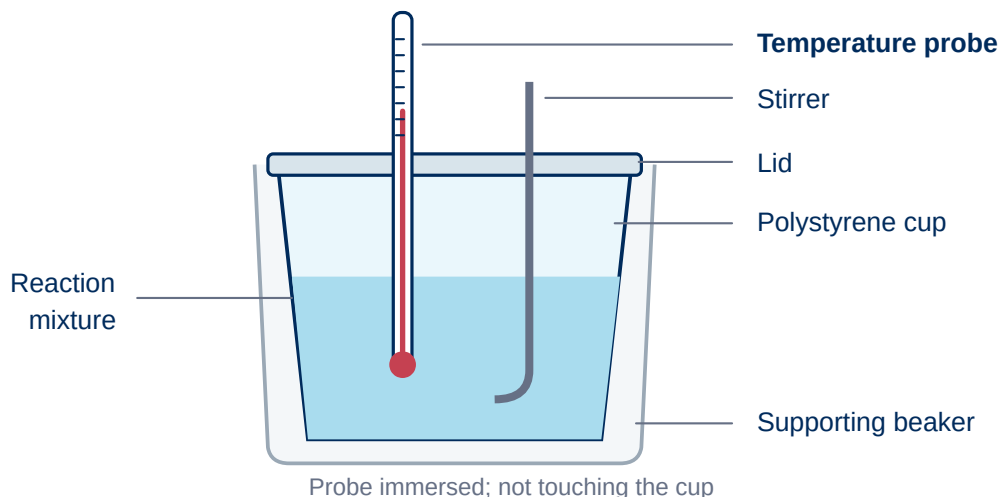
11. Plan a solution calorimetry experiment



SPEC 3.1.4.2; RP2 Describe a workable method for AQA Required practical 2 and explain each key choice.



CRACK IT Measure a temperature change in an insulated cup, then connect that heat change to the amount that reacted.



1. Measure a known volume of a solution of known concentration into a polystyrene cup supported in a beaker. Fit a lid with a hole for a temperature probe.
2. Record temperature at fixed time intervals before mixing. Bring both solutions to the same initial temperature, or account for different starting temperatures.
3. At a recorded time, add the measured second reactant promptly, replace the lid and stir gently. Keep the probe immersed without touching the cup.
4. Continue recording regularly as temperature rises or falls and then returns towards room temperature. Use the trace to estimate the corrected change.
5. Use the combined solution mass and supplied c for q . Use the balanced equation and limiting reagent for n . Repeat comparable runs and compare the results.

Worked choice: $25.0 \text{ cm}^3 \text{ HCl} + 25.0 \text{ cm}^3 \text{ NaOH}$

A polystyrene cup reduces heat transfer; the lid reduces heat transfer and evaporation; gentle stirring makes the temperature more uniform. A measuring cylinder can measure these volumes, but apparatus choice should reflect the required precision. Wear eye protection and use the school's risk assessment for the selected concentrations.



TEST IT 11

- a) Explain why a solution calorimeter uses a polystyrene cup, a lid and gentle stirring. Give one reason for each. [3 marks]

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- b) In a displacement experiment, why record temperature before adding the metal and continue after the maximum? [2 marks]

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✓ **DONE WHEN** I can describe a complete method and justify the apparatus choices.

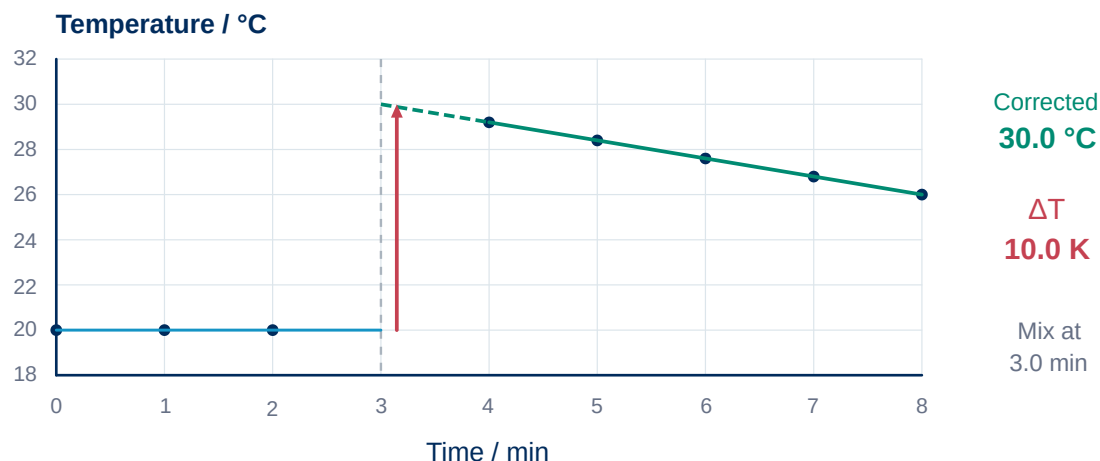
12. Correct a temperature change from a graph



SPEC 3.1.4.2; RP2 Use temperature-time data to estimate the change at the mixing time.



CRACK IT Extend the stable cooling trend back to the mixing time. Measure the gap between the two extrapolated temperatures there.



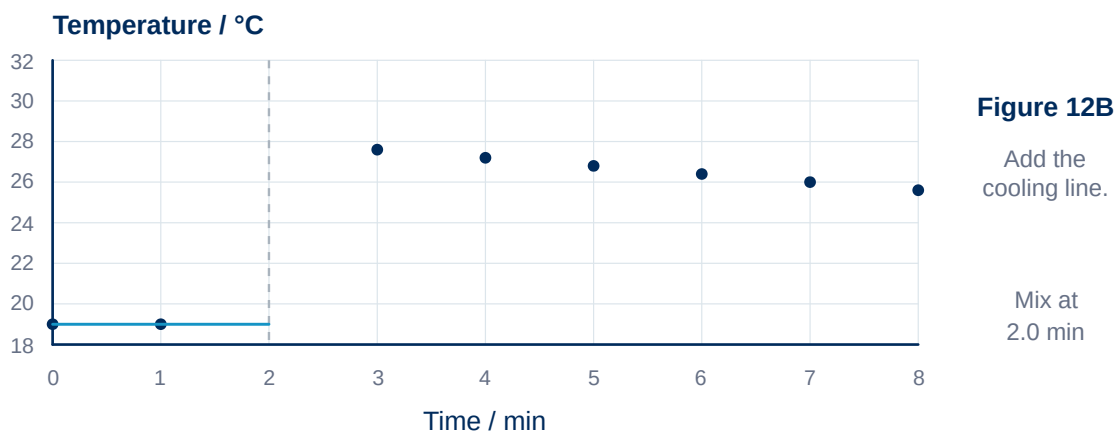
Add the second reactant at 3.0 min. The measured peak is 29.2 °C, but cooling has already occurred. Extrapolating the later straight cooling trend back to 3.0 min gives 30.0 °C. The initial trend gives 20.0 °C there, so corrected $\Delta T = 10.0$ K.

USE THE RIGHT LINE Fit the approximately straight cooling section after the reaction is complete. Do not join every point or extrapolate the rapid warming section. This correction estimates heat exchange; it does not fix incomplete reaction or the calorimeter's heat capacity.



TEST IT 12

a) In Figure 12B, mixing occurs at 2.0 min. Draw the best-fit straight cooling line through the later points and extend it back to 2.0 min. Read the corrected maximum and calculate the corrected temperature rise. [3 marks]



b) Explain why the corrected temperature rise is larger than the rise to the measured maximum. [1 mark]

✓ **DONE WHEN** I can read the two temperatures at the same mixing time.

13. Explain errors and useful improvements



SPEC 3.1.4.2; RP2 Connect a specific experimental limitation to the direction of the enthalpy error.



CRACK IT Name the problem, say which measurement it changes, then explain the effect on the calculated magnitude of ΔH .

LIMITATION	EFFECT ON THE RESULT	USEFUL IMPROVEMENT
Heat escapes from a hot solution	Measured rise and heat gained are too small; exothermic ΔH is less negative.	Insulate, use a lid and correct with a temperature-time trace.
Apparatus absorbs heat but its heat capacity is ignored	Heat calculated from water alone is too small.	Include the calorimeter heat capacity if supplied or calibrated.
Fuel burns incompletely	Less heat released per mole of fuel than for complete combustion.	Improve oxygen access and check for soot.
Fuel evaporates between weighings	Apparent mass burned and n are too large; the magnitude of ΔH is too small.	Cap the burner when extinguished; weigh promptly.
Random reading variation	Repeated values scatter.	Repeat comparable runs and calculate a justified mean.

Worked explanation: combustion result

Measured $\Delta H = -850 \text{ kJ mol}^{-1}$; accepted $\Delta H = -1100 \text{ kJ mol}^{-1}$. Some heat warms the room rather than the water. The measured ΔT and q are too small, so division by the fuel moles gives a value with too small a magnitude: less negative.

For an endothermic dissolution, heat entering from the room reduces the observed temperature fall. The calculated positive ΔH is then too small. Repeats alone do not remove this systematic heat-transfer error.

PRECISE LANGUAGE "Human error" is not an explanation. Give a cause and direction. Keep "less negative" separate from "smaller magnitude".



TEST IT 13

a) An endothermic dissolution gives $+12.0 \text{ kJ mol}^{-1}$ instead of $+18.0 \text{ kJ mol}^{-1}$. Explain how heat exchange with the room could cause this. [3 marks]

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b) Explain why repeating a combustion experiment and averaging does not remove heat loss. State one improvement directed at that error. [2 marks]

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✓ **DONE WHEN** I can explain the direction of an error and match it to an improvement.

14. Handle uncertainty and significant figures



SPEC 3.1.4.2; MS 1.1; practical skills Calculate measurement uncertainty and report a defensible precision.



CRACK IT A temperature change or mass loss is a difference of two readings. Include the uncertainty from both.

UNCERTAINTY Measurement uncertainty estimates a range around a measured value. Absolute uncertainty uses the same units as the measurement; percentage uncertainty expresses that range relative to the measured value.

$$\text{Percentage uncertainty} = (\text{absolute uncertainty} \div \text{measured value}) \times 100$$

Worked example 1: a temperature difference

Each temperature reading has uncertainty $\pm 0.1^\circ\text{C}$. A rise from 21.4°C to 26.8°C is 5.4 K . Using maximum uncertainties, uncertainty in $\Delta T = \pm(0.1 + 0.1) = \pm 0.2\text{ K}$. Percentage uncertainty = $0.2 \div 5.4 \times 100 = 3.7\%$.

Worked example 2: burner mass loss

Each balance reading has uncertainty $\pm 0.01\text{ g}$. A mass loss of 0.80 g has absolute uncertainty $\pm 0.02\text{ g}$ and percentage uncertainty $0.02 \div 0.80 \times 100 = 2.5\%$. A larger mass loss reduces this percentage if the balance uncertainty stays unchanged.

RECORD AND ROUND Record raw readings to the instrument's resolution. Keep extra digits in intermediate calculations. Round the final result to a precision justified by the data; a calculator's long display does not make the experiment more accurate.

Worked example 3: round the final value

For multiplication and division, the measured input with the fewest significant figures usually limits the final precision. If all limiting inputs have three significant figures, $-57.236\text{ kJ mol}^{-1}$ is reported as -57.2 kJ mol^{-1} .

If a question requests a combined maximum percentage uncertainty for quantities multiplied or divided, add the relevant percentage uncertainties. Treat supplied constants as exact unless their uncertainties are given. This estimate does not include systematic heat losses.



TEST IT 14

a) A thermometer has uncertainty $\pm 0.2^\circ\text{C}$ per reading. The temperature changes from 18.4°C to 26.4°C . Calculate the maximum absolute and percentage uncertainty in the temperature rise. [3 marks]

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b) A burner loses 0.50 g . Each weighing has uncertainty $\pm 0.01\text{ g}$. Calculate the percentage uncertainty in mass loss and explain why a larger loss would reduce it. [3 marks]

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c) A calculation gives $\Delta H = -41.6789\text{ kJ mol}^{-1}$ from measurements given to three significant figures. Report the result to an appropriate precision. [1 mark]

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✓ **DONE WHEN** I can include both readings and distinguish uncertainty from systematic error.

15. Work backwards to an unknown



SPEC 3.1.4.2 Rearrange calorimetry equations to find a temperature change or a reacting amount.



CRACK IT Write the equation containing the unknown, rearrange it, then calculate any missing quantity using a second equation.

Worked example 1: predict a temperature rise

0.0200 mol of reaction releases 55.0 kJ per mole. Assume all its heat warms 80.0 g of solution with $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$. Find the temperature rise.

1. The unknown is ΔT : $q = mc\Delta T$, so $\Delta T = q \div (mc)$.
2. Find the missing q : $q(\text{solution}) = -n\Delta H = -0.0200 \times (-55.0) = +1.10 \text{ kJ} = 1100 \text{ J}$.
3. Substitute: $\Delta T = 1100 \div (80.0 \times 4.18) = 3.29 \text{ K}$.

Worked example 2: find a reacting amount

A process has $\Delta H = -40.0 \text{ kJ mol}^{-1}$ and releases 2.00 kJ. $\Delta H = q(\text{reaction}) \div n$, so $n = q(\text{reaction}) \div \Delta H = (-2.00) \div (-40.0) = 0.0500 \text{ mol}$. The amount is positive.

UNKNOWN	REARRANGEMENT	CHECK
ΔT	$q \div (mc)$	q in J if c is $\text{J g}^{-1} \text{ K}^{-1}$
m	$q \div (c\Delta T)$	This is the heated mass
n	$q(\text{reaction}) \div \Delta H$	Use the same energy unit in numerator and denominator



TEST IT 15

a) 0.0150 mol of a reaction has $\Delta H = -60.0 \text{ kJ mol}^{-1}$. Assume all released heat warms 50.0 g of solution with $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$. Calculate the predicted temperature rise. [4 marks]

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b) An endothermic process absorbs 1.80 kJ and has $\Delta H = +30.0 \text{ kJ mol}^{-1}$. Calculate the amount that reacts. [2 marks]

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✓ **DONE WHEN** I can begin with the unknown and build a linked calculation route.

16. Mixed practice and final checks



SPEC 3.1.4.1–3.1.4.2 Combine definitions, calorimetry, mole ratios and practical reasoning.



CRACK IT Identify the heated material, calculate its heat change, find the reacting amount and check the sign and units.

Worked checking example

A student calculates $q = +1800 \text{ J}$ for a solution that warms, then writes $\Delta H = +1800 \div 0.0300 = +60\,000 \text{ kJ mol}^{-1}$. There are two errors: reverse the sign for the reaction and convert J to kJ. Correct result: $\Delta H = -1.800 \div 0.0300 = -60.0 \text{ kJ mol}^{-1}$.

FINAL CHECK Can you define formation and combustion, label an enthalpy profile, identify both masses, apply a mole ratio, correct ΔT and explain a practical error? Use the answers to check each step of your method.



TEST IT 16

a) Write the standard formation equation for NaCl(s) , including state symbols. [2 marks]

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b) 50.0 cm^3 of $0.100 \text{ mol dm}^{-3} \text{ HCl}$ is mixed with 25.0 cm^3 of $0.100 \text{ mol dm}^{-3} \text{ Ba(OH)}_2$. The corrected temperature rise is 0.900 K . Using $2\text{HCl} + \text{Ba(OH)}_2 \rightarrow \text{BaCl}_2 + 2\text{H}_2\text{O}$, calculate ΔH per mole of water formed. Use density = 1.00 g cm^{-3} and $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$. [6 marks]

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c) For a hot solution, explain how ignoring heat absorbed by the cup affects the calculated exothermic ΔH . [2 marks]

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✓ **DONE WHEN** I can solve an unfamiliar calculation and justify the experimental assumptions.

Answers for TEST IT 1 to 6

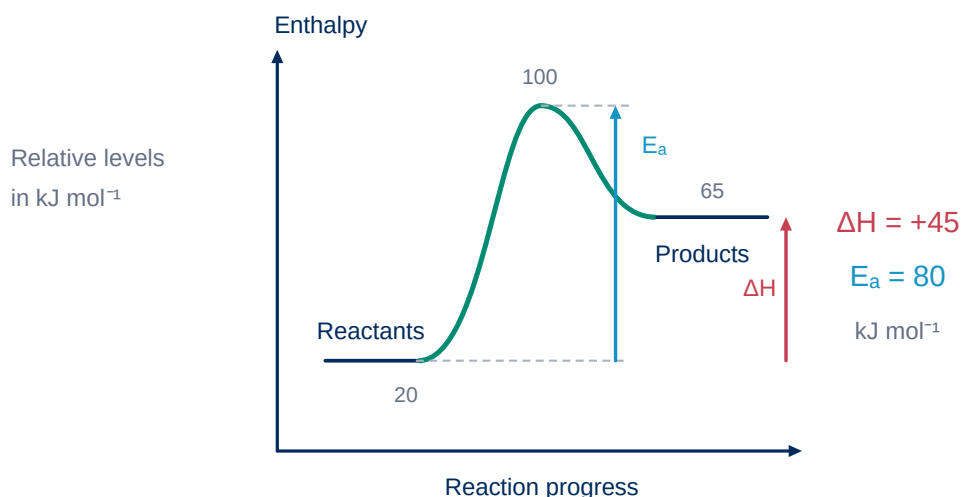
Use these complete worked answers to check both the reasoning and the final result. Mark allocations refer to these original questions.

TEST IT 1 | 5 marks

- a) Exothermic; heat transfers from the reacting system to the surroundings; ΔH is negative.
 b) Higher. The reacting system absorbs heat from the surroundings, so its enthalpy increases.

TEST IT 2 | 5 marks

- a) $\Delta H = 65 - 20 = +45 \text{ kJ mol}^{-1}$; $E_a = 100 - 20 = 80 \text{ kJ mol}^{-1}$. Endothermic profile with products above reactants; arrows shown below.



- b) Particles must overcome the activation energy barrier before reaction can occur.

TEST IT 3 | 5 marks

- a) 100 kPa; 298 K (25 °C); 1 mol dm⁻³.
 b) The fuel starts in different physical states. Vaporising liquid ethanol requires energy, so the reaction enthalpy differs even when the products are identical.

TEST IT 4 | 6 marks

- a) Enthalpy change when one mole of compound forms from its constituent elements; all substances in standard states under standard conditions.
 b) $2\text{C(s, graphite)} + 3\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{C}_2\text{H}_5\text{OH(l)}$.
 c) 0 kJ mol⁻¹.

TEST IT 5 | 6 marks

- a) Enthalpy change when one mole of substance burns completely in oxygen; reactants and products in standard states under standard conditions.
 b) $\text{C}_4\text{H}_{10}(\text{g}) + 6\frac{1}{2}\text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + 5\text{H}_2\text{O(l)}$.
 c) The enthalpy is defined for combustion of one mole of butane.

TEST IT 6 | 5 marks

- a) $\Delta T = 7.50 \text{ K}$. $q = 90.0 \times 4.18 \times 7.50 = +2821.5 \text{ J} = +2.82 \times 10^3 \text{ J} = +2.82 \text{ kJ}$.
 b) 50.0 g; both solutions change temperature, so use their combined mass.

Answers for TEST IT 7 to 11

TEST IT 7 | 6 marks

a) $q(\text{solution}) = 100.0 \times 4.18 \times (-2.50) = -1045 \text{ J}$.
 $q(\text{reaction}) = +1.045 \text{ kJ}$.
 $n = 3.00/101.1 = 0.0296736 \text{ mol}$.
 $\Delta H = 1.045/0.0296736 = +35.2 \text{ kJ mol}^{-1}$.

b) The dissolving process absorbs heat from the solution/surroundings, so it is endothermic.

TEST IT 8 | 5 marks

a) $n(\text{HNO}_3) = 0.0200 \text{ mol}$;
 $n(\text{KOH}) = 0.0150 \text{ mol}$, so KOH limits and $n(\text{H}_2\text{O}) = 0.0150 \text{ mol}$.
 $m = 70.0 \text{ g}$;
 $q = 70.0 \times 4.18 \times 2.90 = 848.54 \text{ J}$.
 $\Delta H = -0.84854/0.0150 = -56.6 \text{ kJ mol}^{-1}$.

MOLE CHECK	HNO ₃	KOH	H ₂ O
Coefficient	1	1	1
Initial amount / mol	0.0200	0.0150	—
Amount reacting or formed / mol	0.0150	0.0150	0.0150

TEST IT 9 | 5 marks

a) $n(\text{CuSO}_4) = 0.250 \times 40.0/1000 = 0.0100 \text{ mol}$.
 $q = 40.0 \times 4.18 \times 25.0 = 4180 \text{ J}$.
 $\Delta H = -4.180/0.0100 = -418 \text{ kJ mol}^{-1}$.

b) 85.0 g.

TEST IT 10 | 6 marks

a) Fuel burned = 0.75 g; $\Delta T = 24.00 \text{ K}$.
 $q(\text{water}) = 150.0 \times 4.18 \times 24.00 = 15048 \text{ J}$.
 $n = 0.75/60.0 = 0.0125 \text{ mol}$.
 $\Delta H = -15.048/0.0125 = -1203.84 \text{ kJ mol}^{-1}$, approximately $-1.2 \times 10^3 \text{ kJ mol}^{-1}$ (2 s.f., limited by fuel mass).

b) It is the water whose temperature change and specific heat capacity are being used.

TEST IT 11 | 5 marks

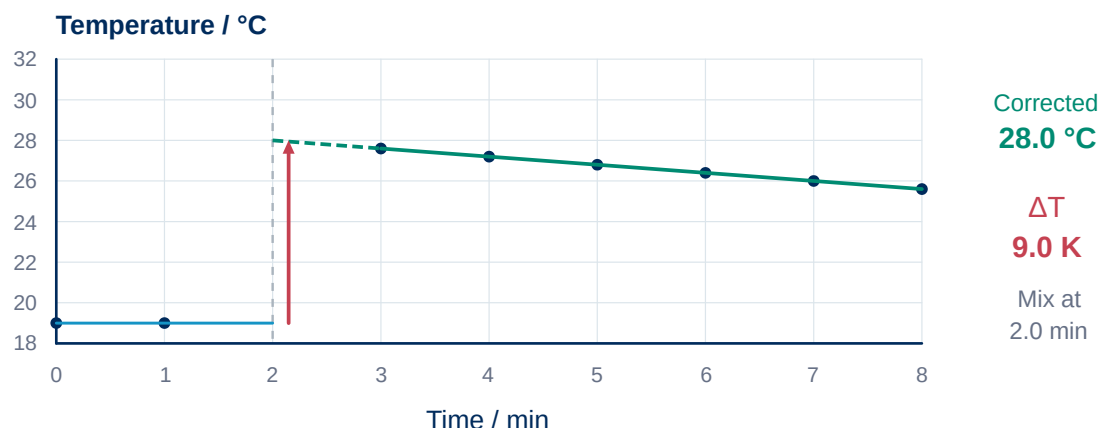
a) Cup: reduces heat transfer to the surroundings. Lid: reduces heat loss/evaporation. Stirring: makes temperature more uniform.

b) The initial readings establish the starting-temperature trend. Later readings establish the cooling trend for extrapolation back to the mixing time.

Answers for TEST IT 12 to 16

TEST IT 12 | 4 marks

a) The cooling line reaches 28.0 °C at 2.0 min. Initial temperature there is 19.0 °C. Corrected $\Delta T = 9.0$ K.



b) Heat was already being lost during the reaction and before the measured maximum was recorded.

TEST IT 13 | 5 marks

a) Heat enters from the warmer room; the observed temperature fall and magnitude of q are too small; the calculated positive ΔH is too small.

b) Heat loss is a systematic limitation affecting each run, so averaging does not remove it. Use insulation/a draught shield to reduce it (while allowing oxygen access).

TEST IT 14 | 7 marks

a) $\Delta T = 8.0$ K; maximum absolute uncertainty = ± 0.4 K. Percentage uncertainty = $0.4/8.0 \times 100 = 5.0\%$.

b) Absolute uncertainty = ± 0.02 g. Percentage = $0.02/0.50 \times 100 = 4.0\%$. A larger measured loss gives a smaller percentage for the same absolute uncertainty.

c) $-41.7 \text{ kJ mol}^{-1}$ (three significant figures).

TEST IT 15 | 6 marks

a) $q(\text{solution}) = -0.0150 \times (-60.0) = 0.900 \text{ kJ} = 900 \text{ J}$.
 $\Delta T = q/(mc) = 900/(50.0 \times 4.18) = 4.31 \text{ K}$.

b) $n = q(\text{reaction})/\Delta H = +1.80/+30.0 = 0.0600 \text{ mol}$.

TEST IT 16 | 10 marks

a) $\text{Na(s)} + \frac{1}{2}\text{Cl}_2(\text{g}) \rightarrow \text{NaCl(s)}$.

b) $n(\text{HCl}) = 0.00500 \text{ mol}$; $n[\text{Ba(OH)}_2] = 0.00250 \text{ mol}$.

The 2:1 amounts react exactly;

$n(\text{H}_2\text{O}) = 0.00500 \text{ mol}$.

$m = 75.0 \text{ g}$;

$q = 75.0 \times 4.18 \times 0.900 = 282.15 \text{ J}$.

$\Delta H = -0.28215/0.00500 = -56.4 \text{ kJ mol}^{-1}$.

c) Some released heat warms the cup rather than the solution. Calculating q for the solution alone underestimates the released energy, so ΔH is less negative.

Specification and practical sources

AQA Chemistry 7405, 3.1.4.1–3.1.4.2 and Required practical 2. Activation energy is a supporting link to 3.1.5.1. Hess's law and mean bond enthalpies are reserved for Energetics II.

Official specification: [aqa.org.uk/subjects/chemistry/a-level/chemistry-7405/specification/subject-content/physical-chemistry](https://www.aqa.org.uk/subjects/chemistry/a-level/chemistry-7405/specification/subject-content/physical-chemistry)

AQA Practical 2 apparatus guide: filestore.aqa.org.uk/resources/chemistry/AQA-7404-7405-P2.PDF