Energetics I

OCR A Chemistry A H432 | Specification-led student notes

SPEC → CRACK IT → WORKED EXAMPLE → TEST IT → DONE

SPECIFICATION COVERAGE

OCR A REF	THIS PACK TEACHES	SECTIONS
3.2.1(a)–(c)	Energy transfer, signs, enthalpy profiles and activation energy	1–2
3.2.1(d)	Standard conditions, states and enthalpy definitions	3–4
3.2.1(e); (h), direct methods	Calorimetry, molar enthalpies and neutralisation	5–8
3.2.1(e); (h), direct methods	Combustion measurements and calculations	9–10
3.2.1(e), (h); practical skills	Cooling curves, limitations, improvements and uncertainty	11–12

YOUR AIM Turn a measured temperature change into a justified molar enthalpy change, with the correct mass, mole quantity, sign and unit.

LESSON BOUNDARY Covers OCR 3.2.1(a)–(e) and direct determination from (h), including PAG3 skills. Hess cycles, indirect determination and average bond-enthalpy calculations belong to Energetics II.

OCR A Chemistry A H432 | Specification version 3.1 (May 2026)

Keep three quantities separate

Temperature change, ΔT, describes what the thermometer measures. Energy transfer, q, describes the energy gained or lost by the water or solution. Molar enthalpy change, ΔH, describes the reaction per mole on the stated basis.

How to use these notes

Write the equation containing the unknown first. Find any missing quantities, then substitute and check the sign. Use the completed grids to follow mole ratios; attempt the TEST IT questions before reading the answers.

Original CrackingChem questions with suggested self-checking marks. Unless stated otherwise, give calculated molar enthalpies to three significant figures.

1. Track energy between system and surroundings

SPEC 3.2.1(a) Explain exothermic and endothermic reactions and the sign of ΔH.

CRACK IT The system is the reacting chemicals. Energy passes between the system and the surroundings, which include the water or solution whose temperature you measure.

An enthalpy change is the heat energy change at constant pressure. For a reaction, compare the enthalpy of the products with that of the reactants.

$$\Delta H = H(\text{products}) - H(\text{reactants})$$

FEATURE	EXOTHERMIC	ENDOTHERMIC
Energy transfer	System → surroundings	Surroundings → system
Surroundings temperature	Rises	Falls
Relative product enthalpy	Lower than reactants	Higher than reactants
Sign of ΔH	Negative	Positive

WORKED EXAMPLE | A temperature rise

A reaction in a cup warms the solution from 20.0 °C to 26.0 °C.

The solution has gained energy. That energy came from the reacting system, so the reaction is exothermic and ΔH is negative.

The temperature rise is positive, but the reaction's enthalpy change is negative.

EXAM CHECK Do not decide the sign from "energy change" alone. State whose energy changes. If the surroundings gain 500 J, the reacting system loses 500 J, under the usual calorimetry assumptions.

TEST IT 1 A reaction takes in 1.25 kJ from its surroundings.
(a) State whether it is exothermic or endothermic. [1]
(b) State the sign of ΔH. [1]
(c) State how the surroundings' temperature changes. [1]
(d) Compare the enthalpy of the products with that of the reactants. [1]

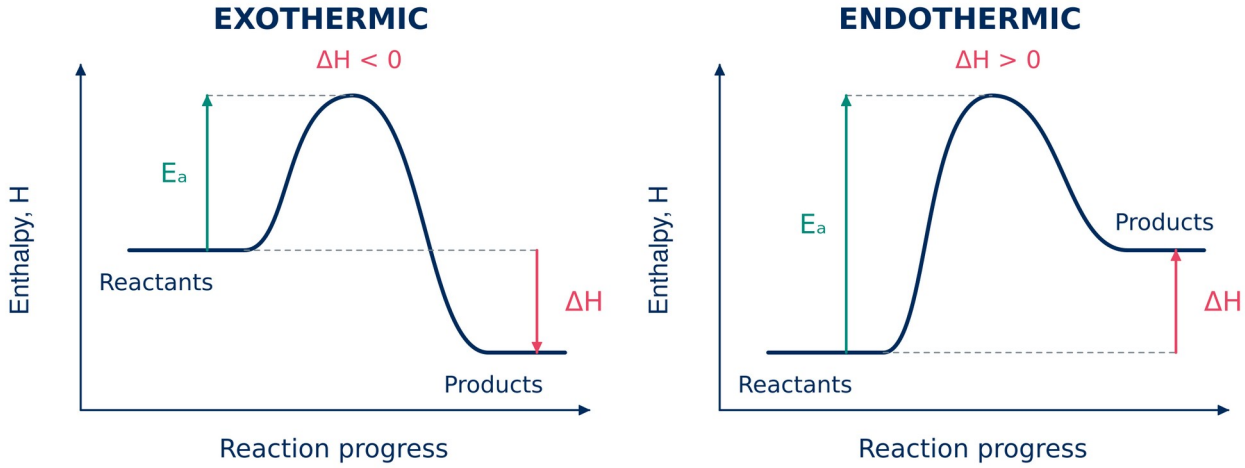
[4 marks]

✓ DONE WHEN You can connect energy transfer, temperature change, relative enthalpy and the sign of ΔH.

2. Draw and read enthalpy profiles

SPEC 3.2.1(b), (c) Construct enthalpy profile diagrams and explain activation energy.

CRACK IT Activation energy, E_a, is the minimum energy required for a reaction to take place. On a profile, measure E_a from the reactant level to the top of the curve.



Label the vertical axis enthalpy and the horizontal axis reaction progress. Draw horizontal reactant and product levels. The ΔH arrow goes from the reactant level to the product level; E_a goes upwards to the peak.

WORKED EXAMPLE | Reading the energy levels

Relative enthalpies are: reactants 80, products 30 and the peak 140 kJ mol^{-1}.

$$\Delta H = 30 - 80 = -50 \text{ kJ mol}^{-1}$$

$$E_a = 140 - 80 = 60 \text{ kJ mol}^{-1}$$

The reaction is exothermic. It still needs activation energy to begin.

EXAM CHECK The horizontal axis is reaction progress, not time. ΔH is the difference between the two levels, not the height of the peak. E_a must start at the reactant level, not the axis baseline.

TEST IT 2 Relative enthalpies are: reactants 25, products 70 and the peak 150 kJ mol^{-1}.
(a) Calculate ΔH and E_a. [2]
(b) Sketch a labelled enthalpy profile, showing both quantities with correctly directed arrows. [3]

[5 marks]

✓ DONE WHEN Your levels, arrow directions and activation-energy measurement agree with the reaction's sign.

3. Standard conditions and the stated equation

SPEC 3.2.1(d)(i), (ii) Use standard conditions, standard states and enthalpy change of reaction.

CRACK IT OCR standard conditions are 100 kPa and a stated temperature, usually 298 K. Standard states are the physical states of substances under those conditions.

298 K is approximately 25 °C. Pure substances are in their standard states; for solutions, the standard concentration is 1 mol dm^{-3}. The superscript ° indicates a standard enthalpy change.

AT 298 K AND 100 kPa	STANDARD STATE
Oxygen	$O_2(g)$
Bromine	$Br_2(l)$
Water	$H_2O(l)$
Carbon	C(s, graphite)

Enthalpy change of reaction

$\Delta_r H$ is the enthalpy change associated with a stated equation. The coefficients tell you the reacting amounts to which the value refers. Always keep the equation and the enthalpy value together.

WORKED EXAMPLE | The equation fixes the amounts

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l) \quad \Delta_r H^\circ = -572 \text{ kJ mol}^{-1}$$

As written, the reaction releases 572 kJ when 2 mol of H_2 reacts and 2 mol of H_2O forms.

For $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$, the corresponding value is −286 kJ mol^{-1}.

Halving every coefficient halves the associated enthalpy change.

EXAM CHECK Standard conditions are not the same as “room temperature and pressure” used with the 24.0 dm^3 mol^{-1} gas-volume approximation. State 100 kPa and the stated temperature when asked about standard enthalpy conditions.

TEST IT 3 (a) State the standard pressure and the temperature normally used for tabulated standard enthalpies. [2]

(b) Give the standard state of bromine at this temperature. [1]

(c) $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ has $\Delta_r H^\circ = -92.0$ kJ mol^{-1}. What is the enthalpy change for forming 1 mol of NH_3 from its elements? [1]

[4 marks]

✓ DONE WHEN You can state the conditions, choose the correct state symbols and link $\Delta_r H$ to the equation as written.

4. Learn the three required definitions

SPEC 3.2.1(d)(iii)–(v) Define and use enthalpy changes of formation, combustion and neutralisation.

CRACK IT Every definition has a one-mole basis. Identify exactly what that one mole refers to before writing the equation or doing a calculation.

STANDARD ENTHALPY CHANGE OF FORMATION | $\Delta_f H°$

The enthalpy change when one mole of a compound is formed from its constituent elements in their standard states, under standard conditions.

$$C(s, \text{graphite}) + 2H_2(g) \rightarrow CH_4(g)$$

Make exactly 1 mol of the compound. Start with elements, not other compounds.

STANDARD ENTHALPY CHANGE OF COMBUSTION | $\Delta_c H°$

The enthalpy change when one mole of a substance is completely burned in oxygen under standard conditions, with all substances in their standard states.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

Burn exactly 1 mol of the substance. Complete combustion of a hydrocarbon forms CO_2 and H_2O.

STANDARD ENTHALPY CHANGE OF NEUTRALISATION | $\Delta_{neut} H°$

The enthalpy change when an acid and an alkali react to form one mole of water under standard conditions, with all substances in their standard states.

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

The one mole refers to water formed, not automatically to acid or alkali used.

EXAM CHECK Fractions in these equations are allowed. If necessary, use $\frac{1}{2}O_2$ rather than doubling the equation and losing the required one-mole basis.

TEST IT 4 (a) Write the equation for the standard enthalpy change of formation of $Fe_2O_3(s)$, including state symbols. [2]
(b) Write the equation for the standard enthalpy change of combustion of ethane, $C_2H_6(g)$, including state symbols. [2]
(c) Define standard enthalpy change of neutralisation. [2]

[6 marks]

✓ **DONE WHEN** You can define the three changes and write equations that match their one-mole bases.

5. Use the correct mass in q = mcΔT

SPEC 3.2.1(e) Calculate energy transfer from a measured temperature change.

CRACK IT q = mcΔT calculates the energy change of the water or solution whose temperature changes. The mass is the mass being heated or cooled.

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

SYMBOL	MEANING AND UNIT
q	Energy gained or lost by the measured water/solution, in J
m	Mass of that water/solution, in g
c	Specific heat capacity, usually 4.18 J g^{-1} K^{-1} for water
ΔT	Temperature change in K; a 1 °C change equals a 1 K change

WORKED EXAMPLE | Two solutions warm up

50.0 cm^3 of one dilute solution is mixed with 25.0 cm^3 of another. The temperature rises by 6.40 K. Assume density 1.00 g cm^{-3} and c = 4.18 J g^{-1} K^{-1}.

Unknown q: use q = mcΔT.

$$m = (50.0 + 25.0) \times 1.00 = 75.0 \text{ g}$$

$$q = 75.0 \times 4.18 \times 6.40 = +2006.4 \text{ J} = +2.01 \text{ kJ}$$

The solution gains energy. Under the calorimetry assumptions, the reaction releases 2.01 kJ.

EXAM CHECK For mixed solutions, use their total mass. OCR also includes an added solid's mass if it fully dissolves or reacts into the solution, unless the question states otherwise. Use the density and heat-capacity assumptions given.

TEST IT 5 2.50 g of a solid dissolves completely in 75.0 g of water. The solution cools by 3.20 K. Use c = 4.18 J g^{-1} K^{-1} and assume no heat exchange with the cup or air.

(a) State the mass to use in q = mcΔT. [1]

(b) Calculate q for the solution, including its sign. [2]

(c) State whether the dissolving process gains or loses energy. [1]

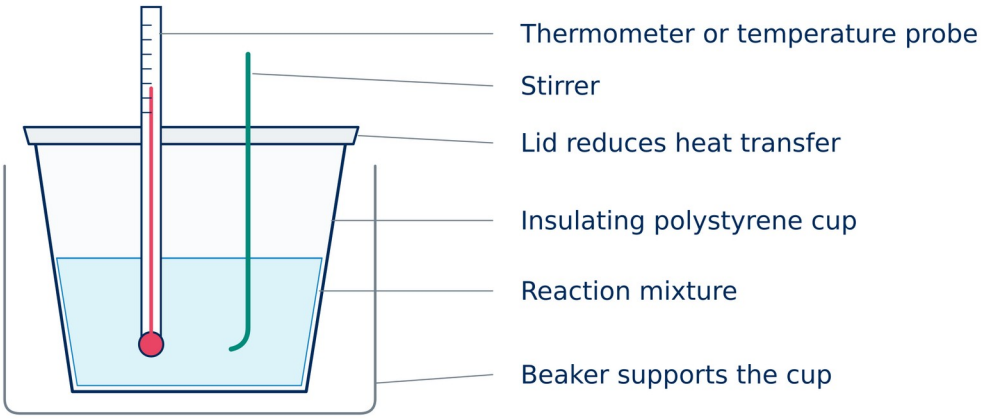
[4 marks]

✓ **DONE WHEN** You can identify the correct mass, retain the sign of ΔT and convert J to kJ.

6. Measure an enthalpy change in a cup

SPEC 3.2.1(e), (h) Direct calorimetry and PAG3 techniques, including neutralisation.

CRACK IT Measure the reacting amounts and the temperature change. Reduce unwanted heat transfer so the measured temperature change reflects the reaction more closely.



1 Measure known volumes and concentrations of acid and alkali. Allow both to reach the same initial temperature and record it.
2 Put one solution in the supported polystyrene cup. Add the other quickly, replace the lid and stir gently.
3 Keep the thermometer bulb or probe immersed without touching the cup. Record temperature regularly before and after mixing.
4 Determine the temperature rise, calculate q for the whole solution and use the reacting amounts to calculate the molar enthalpy change.

FEATURE	PURPOSE
Polystyrene cup and lid	Reduce heat transfer to the wider surroundings.
Gentle stirring	Make the measured temperature representative of the whole solution.
Temperature readings over time	Allow a cooling correction by extrapolation, when required.

EXAM CHECK A temperature rise during neutralisation means the solution gains energy and the reaction is exothermic. Wear eye protection and handle the acid and alkali using the stated risk assessment.

TEST IT 6 Explain why a cup-calorimetry experiment uses:
(a) a lid [1]
(b) stirring [1]
(c) a thermometer bulb immersed in the solution [1]
(d) both solutions at the same starting temperature. [1]

[4 marks]

✓ DONE WHEN You can link each feature of the method to a reliable temperature measurement.

7. Convert q into a molar enthalpy change

SPEC 3.2.1(d)(v), (e) Calculate enthalpy change of neutralisation using the amount of water formed.

CRACK IT For neutralisation, divide the reaction's energy change by moles of water formed. Find the limiting reactant first if one reactant is in excess.

$$\Delta H / \text{kJ mol}^{-1} = \frac{-q / \text{J}}{1000 \times n / \text{mol}}$$

Here q is the signed energy change of the water or solution. A positive q gives a negative reaction enthalpy. This assumes the measured solution accounts for the heat transferred.

For an endothermic reaction, q is negative. If 0.0500 mol reacts while the solution loses 1200 J:

$$\Delta H = -(-1200)/(1000 \times 0.0500) = +24.0 \text{ kJ mol}^{-1}$$

WORKED EXAMPLE | Hydroxide in excess

25.0 cm^3 of 1.00 mol dm^{-3} HCl reacts with 35.0 cm^3 of 1.00 mol dm^{-3} NaOH. The temperature rise is 5.70 K. Use density 1.00 g cm^{-3} and c = 4.18 J g^{-1} K^{-1}.

Start with $\Delta_{neut}H = -q/(1000n(H_2O))$. Find q and $n(H_2O)$.

$$m = 60.0 \text{ g}; q = 60.0 \times 4.18 \times 5.70 = 1429.56 \text{ J}$$

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

MOLE GRID	HCl	NaOH	NaCl	H_2O
Equation ratio	1	1	1	1
n from cV / mol	0.0250	0.0350	—	—
n reacted or formed / mol	0.0250	0.0250	0.0250	0.0250

HCl limits the reaction. Its 1:1 ratio with water gives $n(H_2O)$ = 0.0250 mol. The extra NaOH contributes to the solution mass but does not make extra water.

$$\Delta_{neut}H = -1429.56/(1000 \times 0.0250) = -57.2 \text{ kJ mol}^{-1}$$

TEST IT 7 20.0 cm^3 of 1.00 mol dm^{-3} HNO_3 is mixed with 30.0 cm^3 of 0.800 mol dm^{-3} KOH. The temperature rises by 5.40 K. Calculate the enthalpy change of neutralisation. Identify the limiting reactant in your working. Assume density 1.00 g cm^{-3} and c = 4.18 J g^{-1} K^{-1}.

[5 marks]

✓ DONE WHEN You use total solution mass for q, but only the amount that reacts to determine moles of water formed.

8. Use the equation when the ratio changes

SPEC 3.2.1(d)(v), (e) Apply stoichiometry to neutralisation calculations.

CRACK IT One mole of sulfuric acid can form two moles of water on complete neutralisation. The balanced equation determines the amount of water; the definition determines what you divide by.

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

WORKED EXAMPLE | A 1:2 ratio

20.0 cm³ of 0.500 mol dm⁻³ H_2SO_4 reacts with 40.0 cm³ of 0.500 mol dm⁻³ NaOH. The temperature rise is 4.50 K. Use density 1.00 g cm⁻³ and c = 4.18 J g⁻¹ K⁻¹.

Start with $\Delta_{neut}H = -q/(1000n(H_2O))$.

$$m = 60.0\ g;\ q = 60.0 \times 4.18 \times 4.50 = 1128.6\ J$$

MOLE GRID	H_2SO_4	NaOH	Na_2SO_4	H_2O
Equation ratio	1	2	1	2
n from cV / mol	0.0100	0.0200	—	—
n reacted or formed / mol	0.0100	0.0200	0.0100	0.0200

The amounts are exactly in the 1:2 ratio. The reaction forms 0.0200 mol of water, not 0.0100 mol.

$$\Delta_{neut}H = -1128.6/(1000 \times 0.0200) = -56.4\ kJ\ mol^{-1}$$

EXAM CHECK Dividing by moles of H_2SO_4 here would give the energy change per mole of acid. It would be twice the required neutralisation value per mole of water.

TEST IT 8 25.0 cm³ of 0.600 mol dm⁻³ H_2SO_4 is mixed with 50.0 cm³ of 0.500 mol dm⁻³ NaOH. The temperature rises by 4.50 K.
Identify the limiting reactant, calculate the moles of water formed and determine the enthalpy change of neutralisation. Use density 1.00 g cm⁻³ and c = 4.18 J g⁻¹ K⁻¹.

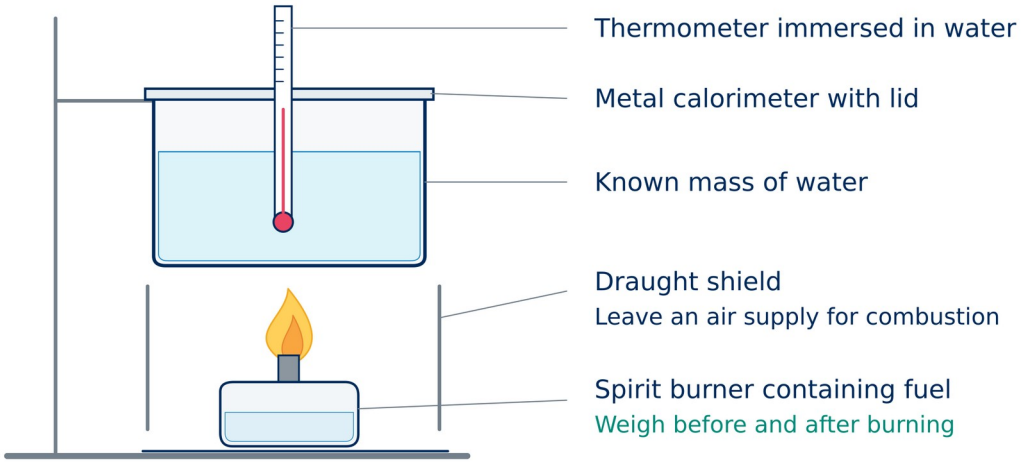
[6 marks]

✓ **DONE WHEN** You compare available moles with the equation ratio and use n(H_2O) in the final division.

9. Measure an enthalpy change of combustion

SPEC 3.2.1(d)(iv), (e), (h) Direct combustion calorimetry and PAG3 techniques.

CRACK IT A burning fuel transfers energy to a known mass of water. Measure both the water's temperature rise and the mass of fuel actually burned.



1 Measure the mass of water in the calorimeter and its initial temperature. Weigh the capped spirit burner with fuel.
2 Remove the cap and light the burner. Heat the water, stirring carefully, with a fixed small gap between flame and calorimeter.
3 Extinguish using the cap after a suitable rise. Record the highest water temperature. Let the capped burner cool and reweigh.
4 Use the decrease in burner mass to find moles of fuel burned. Use the water mass and temperature rise to find q.

Mass of fuel burned = burner mass before − burner mass after

CALCULATION	WHICH MASS?
q = mcΔT	Mass of the water heated
n(fuel) = m(fuel)/M(fuel)	Mass of fuel burned, found by difference

EXAM CHECK Keep spare fuel away from the flame. Never refill a lit or hot burner. A draught shield must still allow an air supply. These open-flame measurements are usually less exothermic than accepted values.

TEST IT 9 (a) State which mass belongs in q = mcΔT in a combustion experiment. [1]
(b) Explain how to obtain the mass used in n = m/M. [1]
(c) Explain why the burner should be capped promptly after extinguishing and kept capped before reweighing. [1]

[3 marks]

✓ DONE WHEN You keep the water mass separate from the fuel mass and can describe the necessary measurements.

10. Calculate the molar enthalpy of combustion

SPEC 3.2.1(d)(iv), (e) Combine water calorimetry with the amount of fuel burned.

CRACK IT Start with $\Delta_c H = -q/(1000n(\text{fuel}))$, with q in J. Find q from the water and n from the fuel, then substitute into the original equation.

WORKED EXAMPLE | Burning ethanol

A burner loses 0.460 g of ethanol, C_2H_5OH. It warms 150.0 g of water by 18.4 K. M(ethanol) = 46.0 g mol^{-1} and c(water) = 4.18 J g^{-1} K^{-1}.

Target: $\Delta_c H = -q/(1000n(\text{fuel}))$.

$$q = mc\Delta T = 150.0 \times 4.18 \times 18.4 = 11\,536.8 \text{ J}$$

$$n(\text{ethanol}) = m/M = 0.460/46.0 = 0.0100 \text{ mol}$$

$$\Delta_c H = -11\,536.8/(1000 \times 0.0100)$$

$$\Delta_c H = -1.15 \times 10^3 \text{ kJ mol}^{-1} \text{ (3 s.f.)}$$

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$

The negative sign shows energy release by combustion. The value is an experimental estimate. Do not label it a standard value unless the appropriate standard-state and condition requirements are met.

EXAM CHECK Do not use the mass of fuel in $q = mc\Delta T$ or divide by moles of water heated. Retain unrounded calculator values until the final result and include kJ mol^{-1}.

TEST IT 10 0.600 g of propan-1-ol, C_3H_7OH, heats 200.0 g of water from 20.00 °C to 38.00 °C. M(propan-1-ol) = 60.0 g mol^{-1} and c(water) = 4.18 J g^{-1} K^{-1}.

(a) Write a complete-combustion equation for 1 mol of propan-1-ol with standard-state symbols. [1]

(b) Calculate its experimental enthalpy change of combustion. [3]

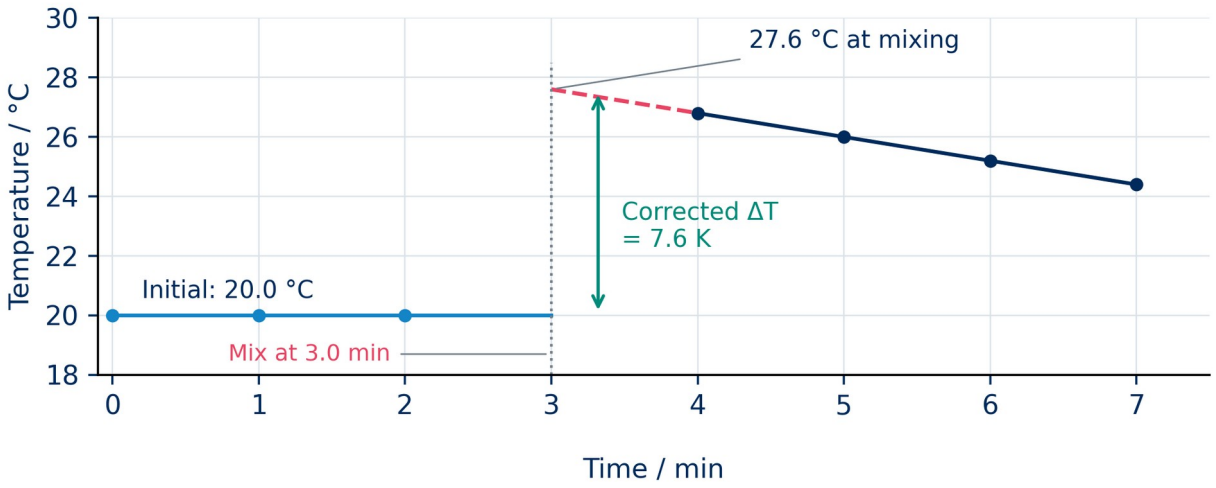
[4 marks]

✓ DONE WHEN You can move from two measured changes to q, then n(fuel), then a signed molar enthalpy.

11. Correct for heat transfer using a cooling curve

SPEC 3.2.1(e), (h); practical graph skills Estimate the temperature change at the mixing time.

CRACK IT Heat transfer begins while the temperature is still changing. Extrapolating the later cooling line back to the mixing time estimates the temperature before that heat loss.



Record a stable initial temperature before mixing. After the reaction finishes, draw a line of best fit through the later cooling readings. Extend it backwards to the mixing time. Measure the vertical difference from the initial-temperature line at that same time.

WORKED EXAMPLE | Use the extrapolated temperature

The graph gives 27.6 °C at the mixing time and an initial temperature of 20.0 °C.

$$\text{Corrected } \Delta T = 27.6 - 20.0 = 7.6 \text{ K}$$

The highest recorded value was 26.8 °C, giving only 6.8 K without correction. The corrected exothermic ΔH would be more negative.

EXAM CHECK Use the later cooling section, not the steep temperature-rise section, for the extrapolation. The correction estimates heat lost to the surroundings; it does not automatically correct heat absorbed by the cup.

TEST IT 11 Reactants are mixed at 3.0 min. The initial-temperature line is 21.3 °C. A straight cooling line passes through 29.0 °C at 4.0 min and 27.8 °C at 6.0 min.

(a) Extrapolate to find the temperature at 3.0 min. [1]

(b) Calculate the corrected temperature rise. [1]

(c) Explain how using the highest recorded temperature instead would affect the calculated exothermic enthalpy change. [2]

[4 marks]

✓ DONE WHEN You can obtain both temperatures at the mixing time and explain the direction of the correction.

12. Evaluate the result and its uncertainty

SPEC 3.2.1(e), (h); practical evaluation Explain limitations, improvements and their effects on calculated ΔH.

CRACK IT Name the cause, trace its effect through q or n, then state what happens to ΔH. For an exothermic value, “less negative” is clearer than “lower”.

LIMITATION	EFFECT ON AN EXOTHERMIC RESULT	IMPROVEMENT
Heat escapes to the air	Measured ΔT and q are too small; ΔH is less negative.	Lid, insulation or an appropriate cooling correction.
Container absorbs heat	Water/solution alone records too little energy; ΔH is less negative.	Allow for container heat capacity when supplied.
Incomplete combustion	Less energy released per mole of fuel; ΔH is less negative.	Improve oxygen supply; avoid soot formation.
Fuel evaporates	Mass loss overstates fuel burned, so n is too large and ΔH is less negative.	Cap promptly and minimise evaporation.

Using water’s heat capacity and density for a solution is an assumption. Its effect depends on the real values, so do not guess a universal error direction. Repeats help assess random variation but do not remove a consistent heat-loss bias.

WORKED EXAMPLE | Uncertainty in a difference

Each temperature reading has uncertainty ±0.10 °C. The measured rise is 6.00 °C.

$$\text{Uncertainty in } \Delta T = 0.10 + 0.10 = \pm 0.20\ ^\circ\text{C}$$

$$\text{Percentage uncertainty} = (0.20/6.00) \times 100 = 3.3\%$$

Two readings contribute to a temperature difference. Apply the same logic to a mass found by difference.

TEST IT 12 An open-flame combustion experiment gives −930 kJ mol^{-1} compared with an accepted value of −1367 kJ mol^{-1}.

(a) Give two distinct experimental causes of the difference and explain the effect of each. [4]

(b) Each thermometer reading has uncertainty ±0.10 °C. Calculate the percentage uncertainty in a temperature rise of 8.00 °C. [2]

[6 marks]

✓ DONE WHEN Your explanation states how the measurement changes and why that makes the molar enthalpy too small or too large in magnitude.

Answers and working

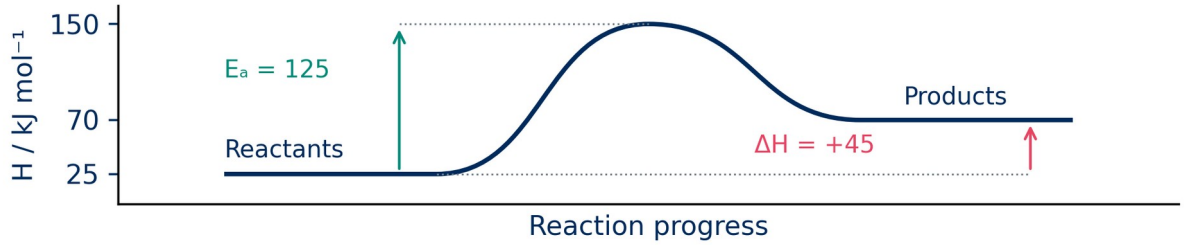
Suggested marking points for the original TEST IT questions. Equivalent chemically correct wording is acceptable.

TEST IT 1 Energy transfer [4 marks]

(a) Endothermic. [1] (b) Positive. [1] (c) The surroundings cool. [1] (d) Products have higher enthalpy than reactants. [1]

TEST IT 2 Enthalpy profile [5 marks]

(a) $\Delta H = 70 - 25 = +45$ kJ mol^{-1} [1]; $E_a = 150 - 25 = 125$ kJ mol^{-1} [1].



(b) Correctly labelled axes and relative reactant/product levels [1]; ΔH arrow upwards from reactants to products [1]; E_a arrow upwards from the reactant level to the peak [1].

TEST IT 3 Standard conditions and amounts [4 marks]

(a) 100 kPa [1] and 298 K (approximately 25 °C) [1].

(b) $Br_2(l)$. [1]

(c) $-92.0/2 = -46.0$ kJ mol^{-1}, for $\frac{1}{2}N_2(g) + 1\frac{1}{2}H_2(g) \rightarrow NH_3(g)$. [1]

TEST IT 4 One-mole definitions [6 marks]

(a) $2Fe(s) + 1\frac{1}{2}O_2(g) \rightarrow Fe_2O_3(s)$. Correct species and coefficients for 1 mol Fe_2O_3 [1]; state symbols [1].

(b) $C_2H_6(g) + 3\frac{1}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$. Correct species and coefficients for 1 mol ethane [1]; state symbols [1].

(c) The enthalpy change when an acid and an alkali react to form one mole of water [1], under standard conditions with substances in their standard states [1].

Answers to solution calorimetry

TEST IT 5 A solution cools [4 marks]

(a) 75.0 + 2.50 = 77.5 g. [1]

(b) ΔT = −3.20 K and q = 77.5 × 4.18 × (−3.20) [1] = −1036.64 J, or −1.04 kJ [1].

(c) The dissolving process gains energy from the solution; it is endothermic. [1]

TEST IT 6 Cup method [4 marks]

(a) A lid reduces heat transfer to the surroundings. [1]

(b) Stirring makes temperature uniform throughout the solution. [1]

(c) The thermometer then measures the solution temperature, rather than that of the air or container. [1]

(d) Otherwise mixing the two solutions itself transfers heat between them, complicating the temperature change attributed to reaction. [1]

TEST IT 7 A 1:1 neutralisation [5 marks]

m = (20.0 + 30.0) × 1.00 = 50.0 g. [1]

q = 50.0 × 4.18 × 5.40 = 1128.6 J. [1]

$n(HNO_3)$ = 0.0200 × 1.00 = 0.0200 mol; n(KOH) = 0.0300 × 0.800 = 0.0240 mol. HNO_3 is limiting. [1]

The 1:1 acid:water ratio gives $n(H_2O)$ = 0.0200 mol. [1]

$\Delta_{neut}H$ = −1128.6/(1000 × 0.0200) = −56.4 kJ mol^{-1}. [1]

TEST IT 8 A 1:2 neutralisation [6 marks]

MOLE GRID	H_2SO_4	NaOH	H_2O
Equation ratio	1	2	2
Available n / mol	0.0150	0.0250	—
n reacted or formed / mol	0.0125	0.0250	0.0250

$n(H_2SO_4)$ = 0.0250 × 0.600 = 0.0150 mol; n(NaOH) = 0.0500 × 0.500 = 0.0250 mol [1]. NaOH is limiting because 0.0150 mol acid would require 0.0300 mol NaOH [1].

$n(H_2O)$ = 0.0250 mol [1]. Total solution mass = 75.0 g [1]. q = 75.0 × 4.18 × 4.50 = 1410.75 J [1]. $\Delta_{neut}H$ = −1410.75/(1000 × 0.0250) = −56.4 kJ mol^{-1} [1].

Answers to combustion and evaluation

TEST IT 9 Two different masses [3 marks]

(a) The mass of water heated. [1]

(b) Subtract the mass of the burner after burning from its mass before burning. [1]

(c) Capping reduces fuel evaporation, which would otherwise be included incorrectly in the mass burned. [1]

TEST IT 10 Combustion calculation [4 marks]

(a) $C_3H_7OH(l) + 4½O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$. [1]

(b) $\Delta T = 38.00 - 20.00 = 18.00$ K; $q = 200.0 \times 4.18 \times 18.00 = 15\,048$ J. [1]

n(propan-1-ol) = 0.600/60.0 = 0.0100 mol. [1]

$\Delta_cH = -15\,048/(1000 \times 0.0100) = -1.50 \times 10^3$ kJ mol^{-1}. [1]

TEST IT 11 Cooling correction [4 marks]

(a) The cooling line falls by 1.2 °C in 2.0 min, or 0.6 °C min^{-1}. Going back from 4.0 to 3.0 min gives 29.6 °C. [1]

(b) Corrected $\Delta T = 29.6 - 21.3 = 8.3$ K. [1]

(c) The highest measured value gives a smaller ΔT and q [1], so the exothermic ΔH is less negative / smaller in magnitude [1].

TEST IT 12 Limitations and uncertainty [6 marks]

(a) Any two distinct valid cause-and-effect pairs, 2 marks each. Examples:

Heat escapes to the air [1], so the water's ΔT and calculated q are too small and ΔH is less negative [1].

Incomplete combustion [1] releases less energy per mole of fuel than complete combustion, making ΔH less negative [1].

Fuel evaporation [1] makes the measured mass loss and n(fuel) too large, making ΔH less negative [1].

(b) Uncertainty in $\Delta T = 0.10 + 0.10 = \pm 0.20$ °C [1]. Percentage uncertainty = (0.20/8.00) × 100 = 2.5% [1].

Specification and practical guidance

Aligned to OCR Chemistry A H432 v3.1 (May 2026), 3.2.1(a)–(e) and direct methods from (h). Practical guidance: OCR Practical Activities Support Guide, PAG3; OCR Chemistry Mathematical Skills Handbook, graph and uncertainty skills. Worked examples and practice questions are original CrackingChem material.

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