



AQA 7405

Energetics I

ENTHALPY CHANGES • CALORIMETRY • REQUIRED PRACTICAL 2

ORIGINAL EXAM QUESTIONS | TEST IT → EXAM STYLE → HARDER MIXED → STRETCH

TEST IT	EXAM STYLE	HARDER MIXED	STRETCH
8	8	4	2

22 QUESTIONS | 207 TOTAL MARKS

Mapped to AQA Chemistry 7405, sections 3.1.4.1-3.1.4.2 and Required Practical 2

Specification checked 17 October 2026

How to use this pack [START HERE](#)

- 1 Read the command word and mark allocation. State is concise; explain must link cause to effect; calculate must show the route and units.
- 2 Use the full-width dotted lines for working. Write equation, rearrangement, substitution and units as separate stages.
- 3 Keep q for the measured surroundings separate from ΔH for the reaction. Preserve unrounded values until the final answer.

Specification route

AQA section	Assessed in this pack
3.1.4.1	Exothermic and endothermic changes; standard conditions; standard enthalpies of formation and combustion.
3.1.4.2 and RP2	$q = mc\Delta T$; molar enthalpy and related calculations; practical methods, correction, uncertainty and evaluation.
Supporting link	A short enthalpy-profile and activation-energy question using 3.1.5.1 language.
Reserved for Energetics II	Hess's law and mean bond enthalpy calculations.

Notation: q in J; m in g; c in $\text{J g}^{-1} \text{K}^{-1}$; ΔT in K; n in mol; ΔH in kJ mol^{-1} .

SECTION A TEST IT

Secure the definitions, signs and calculation discipline before the longer questions.

QUESTION 1 | TEST IT

(a) Define enthalpy change. (2 marks)

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(b) A reaction raises the temperature of its surroundings from 18.6 °C to 24.1 °C. State the type of change, the direction of heat transfer and the sign of ΔH . (3 marks)

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(c) For an endothermic reaction, compare the enthalpy of the products with that of the reactants and explain the difference. (2 marks)

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QUESTION 2 | TEST IT

On one relative energy scale, reactants are at 35 kJ mol^{-1} , products are at -15 kJ mol^{-1} and the peak is at 95 kJ mol^{-1} .

(a) Calculate ΔH for the reaction. (2 marks)

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(b) Calculate the forward activation energy. (2 marks)

.....

(c) Draw the enthalpy profile. Label both axes, reactants, products, ΔH and the forward activation energy. (4 marks)

Draw and label the profile here

**QUESTION 3 | TEST IT**

(a) State the standard pressure, the temperature usually stated and the standard concentration of an aqueous solution. (3 marks)

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(b) Two equations form identical products. One uses $\text{C}_2\text{H}_5\text{OH}(\text{l})$ and the other uses $\text{C}_2\text{H}_5\text{OH}(\text{g})$. Explain why the enthalpy changes differ. (2 marks)

.....

(c) State the standard states at 298 K of carbon and water.

(2 marks)

QUESTION 4 | TEST IT

(a) Define the standard enthalpy of formation.

(3 marks)

(b) Write the standard formation equation for $\text{CaCO}_3(\text{s})$, including state symbols.

(2 marks)

(c) Write the standard formation equation for $\text{NH}_3(\text{g})$, including state symbols.

(2 marks)

(d) State the standard enthalpy of formation of $\text{Cl}_2(\text{g})$.

(1 mark)

QUESTION 5 | TEST IT

(a) Define the standard enthalpy of combustion.

(3 marks)

(b) Write the standard combustion equation for propanone, $\text{C}_3\text{H}_6\text{O}(\text{l})$, including state symbols.

(2 marks)

(c) Explain why the coefficient of propanone must be 1.

(1 mark)

(d) Explain why an experiment that produces carbon monoxide does not measure the standard enthalpy of combustion.

(1 mark)

QUESTION 6 | TEST IT

Use $q = mc\Delta T$ and $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$.

(a) A 125.0 g sample of water warms from 18.35 °C to 26.75 °C. Calculate its heat change in J and kJ, including the sign for the water.

(5 marks)

- (b) A student mixes 35.0 cm^3 and 25.0 cm^3 of two aqueous solutions. Their density is 1.00 g cm^{-3} . State the mass used in $q = mc\Delta T$ and explain why. (2 marks)

QUESTION 7 | TEST IT

A 2.50 g sample of LiCl , $M = 42.4 \text{ g mol}^{-1}$, dissolves completely. The final solution has mass 102.5 g and warms by 7.60 K . Use $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$.

- (a) Calculate the enthalpy change per mole of LiCl . (5 marks)

- (b) Explain the sign of the enthalpy change. (1 mark)

QUESTION 8 | TEST IT

- (a) A thermometer has uncertainty $\pm 0.15^\circ \text{C}$ per reading. The temperature changes from 19.20°C to 26.80°C . Calculate the maximum absolute and percentage uncertainty in ΔT . (3 marks)

- (b) A spirit burner decreases from 112.84 g to 112.19 g . Each balance reading has uncertainty $\pm 0.01 \text{ g}$. Calculate the mass loss, its maximum absolute uncertainty and its percentage uncertainty. (3 marks)

- (c) A calculation based on measurements given to three significant figures gives $\Delta H = -64.285 \text{ kJ mol}^{-1}$. Report the result appropriately. (1 mark)

SECTION B EXAM STYLE

Apply the specification in linked AQA-style calculations, practical work and explanations.

QUESTION 9 | EXAM STYLE

A student dissolves 4.00 g of NH_4NO_3 , $M = 80.0 \text{ g mol}^{-1}$, in 96.0 g of water. The temperature falls from 22.50°C to 18.70°C . Use the final solution mass and $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$.

- (a) State the mass used in $q = mc\Delta T$. (1 mark)

(b) Calculate the heat change of the solution.

(3 marks)

(c) Calculate the enthalpy change per mole of NH_4NO_3 .

(3 marks)

(d) Explain how heat entering from the room affects the calculated value.

(2 marks)

QUESTION 10 | EXAM STYLE

35.0 cm³ of 0.800 mol dm⁻³ HCl is mixed with 25.0 cm³ of 1.00 mol dm⁻³ NaOH. The solutions start at the same temperature and rise by 5.80 K. Use density = 1.00 g cm⁻³, $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$ and $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$.

(a) Calculate the amount of each reactant and identify the limiting reactant.

(3 marks)

(b) Calculate the heat gained by the solution.

(2 marks)

(c) Calculate the enthalpy change per mole of water formed.

(3 marks)

(d) State two assumptions made in this calculation.

(2 marks)

QUESTION 11 | EXAM STYLE

30.0 cm³ of 0.750 mol dm⁻³ H₂SO₄ is mixed with 55.0 cm³ of 0.600 mol dm⁻³ KOH. The corrected temperature rise is 4.20 K. Use density = 1.00 g cm⁻³, $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$ and $\text{H}_2\text{SO}_4 + 2\text{KOH} \rightarrow \text{K}_2\text{SO}_4 + 2\text{H}_2\text{O}$.

(a) Determine the limiting reactant and the amount of water formed.

(4 marks)

(b) Calculate the enthalpy change per mole of water formed.

(3 marks)

(c) Explain why 85.0 g is used as the heated mass.

(1 mark)

(d) The accepted value is $-57.3 \text{ kJ mol}^{-1}$. Explain one experimental reason why the value from this experiment is less negative.

(2 marks)

QUESTION 12 | EXAM STYLE

Excess Fe is added to 75.0 cm^3 of $0.400 \text{ mol dm}^{-3} \text{ CuSO}_4$. The temperature rises by 19.8 K . Use 75.0 g as the heated mass, $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$ and $\text{Fe} + \text{CuSO}_4 \rightarrow \text{FeSO}_4 + \text{Cu}$.

(a) Calculate the enthalpy change per mole of CuSO_4 reacted.

(5 marks)

(b) Explain why the mass of excess iron is not added to the 75.0 g in this model.

(1 mark)

(c) A student adds the whole iron mass to m but still uses the specific heat capacity of water. Explain the effect on the calculated ΔH .

(2 marks)

QUESTION 13 | EXAM STYLE

Figure 1 shows an apparatus used to estimate the enthalpy of combustion of butan-1-ol.

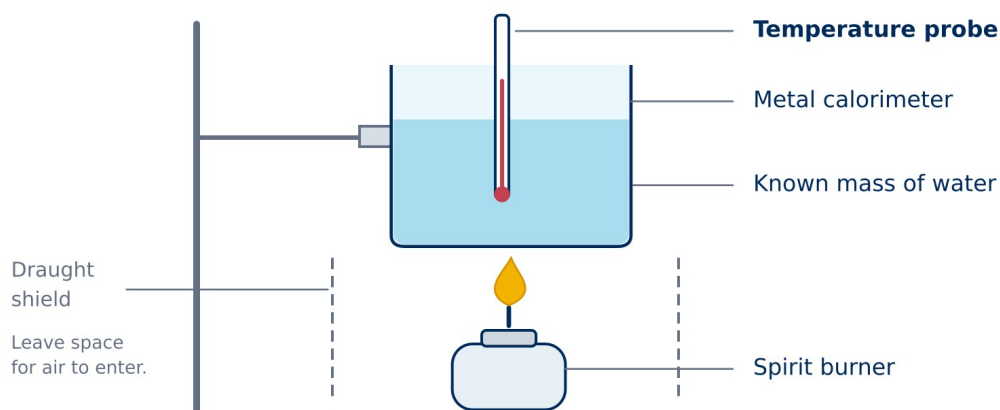


Figure 1 Combustion calorimetry apparatus

The burner decreases from 158.72 g to 157.85 g. A 200.0 g sample of water warms from 17.6 °C to 39.8 °C. Use $M(\text{butan-1-ol}) = 74.0 \text{ g mol}^{-1}$ and $c(\text{water}) = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$.

- (a) Calculate the experimental enthalpy of combustion per mole of butan-1-ol. (6 marks)

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- (b) The accepted value is $-2676 \text{ kJ mol}^{-1}$. Calculate the percentage of the accepted energy captured by the water. (2 marks)

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- (c) Explain how heat transfer to the room affects the experimental enthalpy value. (2 marks)

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- (d) Explain why the burner should be capped and weighed promptly before and after heating. (2 marks)

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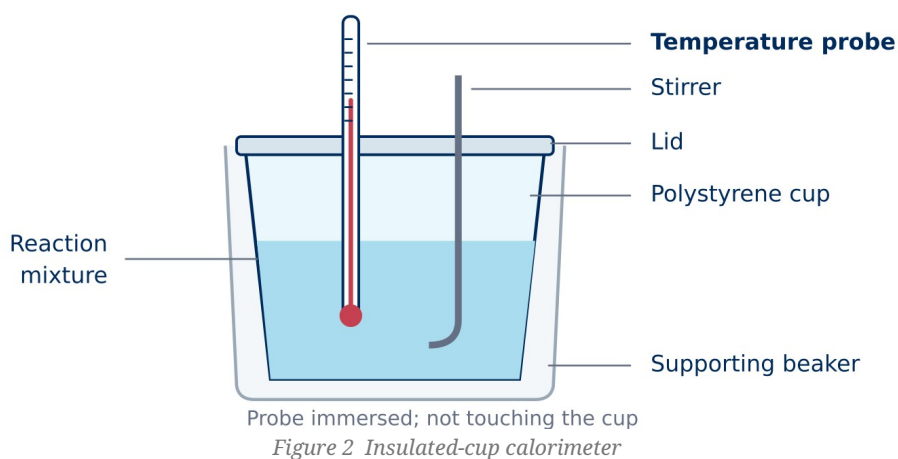
- (e) Explain why the draught shield must still allow air to reach the flame. (1 mark)

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QUESTION 14 | EXAM STYLE

Figure 2 shows the main apparatus for a solution calorimetry experiment.



- (a) Explain the purpose of the polystyrene cup, the lid and gentle stirring. (3 marks)

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- (b) Describe how to obtain a corrected temperature change when measuring the enthalpy change of neutralisation. (5 marks)

- (c) Explain why the temperature probe should not touch the cup. (1 mark)

- (d) State one appropriate safety precaution. (1 mark)

QUESTION 15 | EXAM STYLE

In Figure 3, reactants are mixed at 3.0 min. The initial temperature at that time is 20.0 °C.

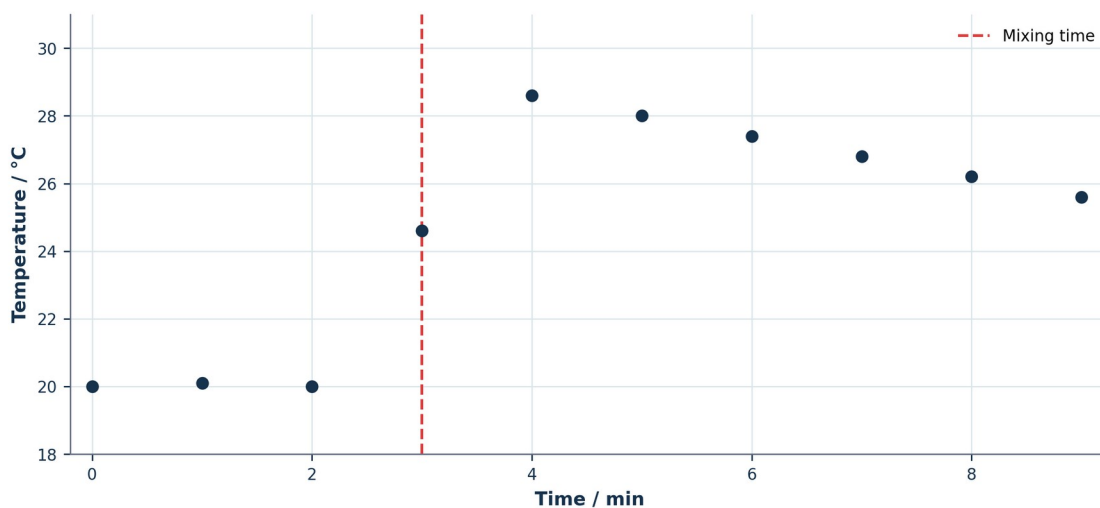


Figure 3 Temperature-time data

- (a) Draw the best-fit straight line through the later cooling points and extrapolate it to 3.0 min. Read the corrected maximum and calculate the corrected temperature rise. (3 marks)

- (b) The measured maximum is 28.6 °C. Calculate the percentage by which using the measured maximum would underestimate the corrected temperature rise. (2 marks)

- (c) Explain why extrapolation gives a larger temperature rise. (2 marks)

QUESTION 16 | EXAM STYLE

A student obtains the following values for the same neutralisation experiment. The accepted value is $-57.3 \text{ kJ mol}^{-1}$.

Trial	Calculated $\Delta H / \text{kJ mol}^{-1}$
1	-52.1
2	-52.3
3	-39.8

(a) Identify the anomalous result and justify your choice. (2 marks)

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(b) Calculate the mean of the concordant results. (1 mark)

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(c) Calculate the percentage difference between this mean and the accepted value. (2 marks)

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(d) Explain why repeating and averaging does not remove systematic heat loss. (2 marks)

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(e) State one improvement directed at reducing heat loss. (1 mark)

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SECTION C HARDER MIXED

Select the route, combine evidence and audit assumptions without student scaffolding.

QUESTION 17 | HARDER MIXED

A 5.00 g sample of Na_2CO_3 , $M = 106.0 \text{ g mol}^{-1}$, dissolves in 95.0 g of water. The temperature rises by 2.95 K . The final solution mass is 100.0 g , $c(\text{solution}) = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$ and the calorimeter heat capacity is 42.0 J K^{-1} .

(a) Calculate the heat gained by the solution and by the calorimeter, then find their total. (4 marks)

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(b) Calculate the enthalpy change per mole of Na_2CO_3 using the total heat gain. (3 marks)

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(c) Calculate the enthalpy change that would be obtained if the calorimeter heat capacity were ignored. (2 marks)

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(d) Calculate the percentage by which ignoring the calorimeter understates the magnitude of ΔH . (2 marks)

(e) Explain the direction of this difference. (1 mark)

QUESTION 18 | HARDER MIXED

50.0 cm³ of HCl of unknown concentration is mixed with 50.0 cm³ of 1.00 mol dm⁻³ NaOH. The temperature rises by 6.20 K. Use density = 1.00 g cm⁻³, $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$ and $\Delta H(\text{neutralisation}) = -57.1 \text{ kJ mol}^{-1}$. Assume all released heat warms the solution.

(a) Calculate the heat gained by the solution. (2 marks)

(b) Use the enthalpy value to calculate the amount of water formed and show that HCl is limiting. (4 marks)

(c) Calculate the concentration of the HCl. (3 marks)

(d) A student incorrectly uses 50.0 g instead of 100.0 g in $q = mc\Delta T$. State the concentration this route would give and explain the direction of the error. (2 marks)

QUESTION 19 | HARDER MIXED

A 2.50 g impure sample of KCl is added to 75.0 g of water. Use 77.5 g as the heated mass, $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$ and $M(\text{KCl}) = 74.6 \text{ g mol}^{-1}$. The temperature falls by 1.40 K. The enthalpy of solution of pure KCl is $+17.2 \text{ kJ mol}^{-1}$. Assume the impurity causes no enthalpy change.

(a) Calculate the heat change of the mixture. (2 marks)

(b) Calculate the amount of pure KCl that dissolved. (2 marks)

(c) Calculate the mass of pure KCl and the percentage purity of the sample. (3 marks)

(d) Explain how heat entering from the room affects the calculated percentage purity. (2 marks)

(e) State the chemical assumption made about the impurity. (1 mark)

QUESTION 20 | HARDER MIXED

A student mixes 20.0 cm³ of 1.50 mol dm⁻³ H₂SO₄ with 50.0 cm³ of 1.00 mol dm⁻³ KOH. The temperature rises by 6.00 K. Use density = 1.00 g cm⁻³, c = 4.18 J g⁻¹ K⁻¹ and H₂SO₄ + 2KOH → K₂SO₄ + 2H₂O.

The student writes: n(H₂SO₄) = 0.0300 mol; n(KOH) = 0.0500 mol; H₂SO₄ limits because 0.0300 is smaller; m = 50.0 g; q = 50.0 × 4.18 × 6.00 = 1254 J; ΔH = +1.254 ÷ 0.0300 = +41.8 kJ mol⁻¹.

(a) Correct the limiting-reagent decision. (2 marks)

(b) Correct the heated mass and calculate q for the solution. (2 marks)

(c) Correct the sign and mole basis for the reaction enthalpy. (2 marks)

(d) Calculate the corrected enthalpy change per mole of water. (2 marks)

(e) The accepted value is -57.3 kJ mol⁻¹. Explain why the experimental value may still be less negative. (2 marks)

SECTION D STRETCH / SYNOPTIC

Connect calorimetry, amount of substance and practical analysis while staying inside the Energetics I boundary.

QUESTION 21 | STRETCH / SYNOPTIC

40.0 cm³ of 0.800 mol dm⁻³ HNO₃ is mixed with 30.0 cm³ of 1.00 mol dm⁻³ Ba(OH)₂ at 2.5 min. Use density = 1.00 g cm⁻³, $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$ and $2\text{HNO}_3 + \text{Ba(OH)}_2 \rightarrow \text{Ba(NO}_3)_2 + 2\text{H}_2\text{O}$. Each temperature reading has uncertainty $\pm 0.1^\circ\text{C}$.

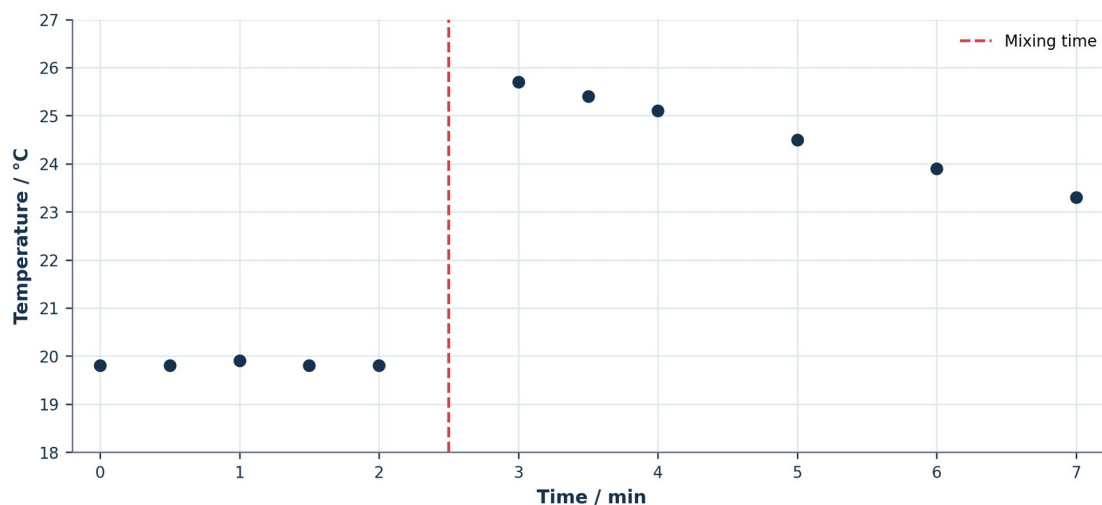


Figure 4 Temperature-time data for the neutralisation

- (a) Extrapolate the cooling line to 2.5 min. State the initial temperature, corrected maximum and corrected ΔT . (4 marks)

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- (b) Determine the limiting reactant and the amount of water formed. (4 marks)

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- (c) Calculate the enthalpy change per mole of water formed. (4 marks)

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- (d) Calculate the maximum percentage uncertainty in the corrected ΔT . (2 marks)

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- (e) The measured maximum is 25.7°C . Calculate the enthalpy value obtained without correction and explain the difference. (2 marks)

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QUESTION 22 | STRETCH / SYNOPTIC

An unknown alcohol is methanol, ethanol or propan-1-ol. A 0.600 g sample burns and heats 150.0 g of water by 12.0 K. Use $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$. Only 55.0% of the energy released is captured by the water.

Candidate	M / g mol ⁻¹	$\Delta_c H^\circ$ / kJ mol ⁻¹
Methanol	32.0	-726
Ethanol	46.0	-1367
Propan-1-ol	60.0	-2021

(a) Calculate the heat captured by the water. (2 marks)

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(b) Calculate the total energy released by the burning sample. (2 marks)

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(c) For each candidate, calculate the energy released by complete combustion of 0.600 g. Deduce the identity of the alcohol. (4 marks)

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(d) For the identified alcohol, predict the water temperature rise if 75.0% of the released energy were captured. (3 marks)

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(e) Explain how evaporation of fuel between weighings would affect the comparison. (3 marks)

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(f) Give two reasons why an open-flame result is an estimate rather than an accurate standard enthalpy of combustion. (2 marks)

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BEFORE YOU TURN OVER - FINAL CHECK

1	For a warmer solution, $q(\text{solution})$ is positive but $q(\text{reaction})$ and ΔH are negative.
2	Use the mass that changes temperature in $q = mc\Delta T$ and the reacting amount in the molar calculation.
3	Use the balanced equation to identify the limiting reactant and the exact mole basis requested.
4	Keep J and kJ consistent, preserve unrounded working and report sensible significant figures.

QUESTIONS END HERE • ANSWERS BEGIN ON THE NEXT PAGE

Energetics I

ANSWERS & MARK SCHEME

FULLY WORKED • PRECISELY MARKED • BUILT FOR IMPROVEMENT

22 ANSWERS | 207 MARKS

SECTION A TEST IT ANSWERS

Use each marking point to identify exactly where a response gained or lost credit.

QUESTION 1 | TEST IT

7 marks

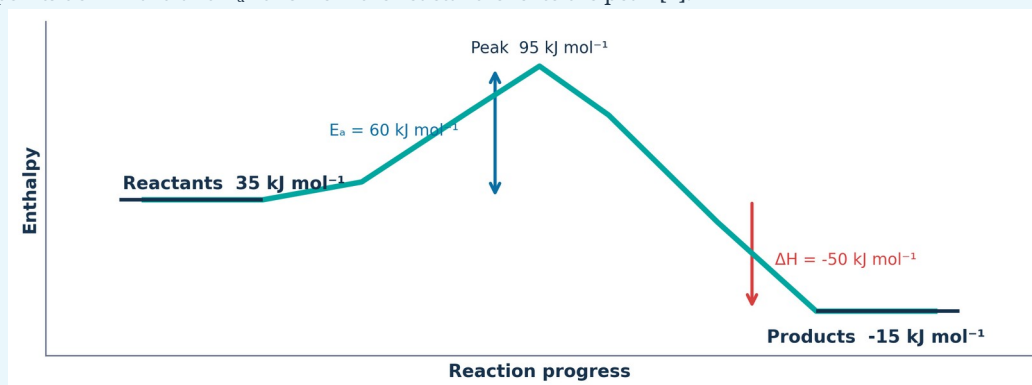
- (a) The heat energy change [1] measured at constant pressure [1].
 (b) The reaction is exothermic [1]. Heat transfers from the reacting system to the surroundings [1], so ΔH is negative [1].
 (c) The products have higher enthalpy than the reactants [1] because the reacting system has absorbed heat from the surroundings [1].

COMMON ERROR A temperature rise in the surroundings means the reaction has released heat. It does not make the reaction ΔH positive.

QUESTION 2 | TEST IT

8 marks

- (a) $\Delta H = H(\text{products}) - H(\text{reactants})$ [1] = $-15 - 35 = -50 \text{ kJ mol}^{-1}$ [1].
 (b) $E_a = 95 - 35$ [1] = 60 kJ mol^{-1} [1].
 (c) Axes labelled enthalpy and reaction progress [1]; reactant and product levels correctly positioned [1]; a curve rises to the stated peak [1]; ΔH points downward and E_a runs from the reactant level to the peak [1].



COMMON ERROR ΔH compares product and reactant levels. Activation energy compares the reactant level with the peak.

QUESTION 3 | TEST IT

7 marks

- (a) 100 kPa [1]; usually 298 K [1]; 1 mol dm^{-3} for an aqueous standard solution [1].
 (b) The ethanol begins in a different physical state [1]. Energy is required to vaporise the liquid, so the two starting enthalpies and reaction enthalpies differ [1].
 (c) Carbon is graphite, C(s, graphite) [1]; water is $\text{H}_2\text{O}(\text{l})$ [1].

QUESTION 4 | TEST IT

8 marks

- (a) The enthalpy change when one mole of a compound is formed [1] from its constituent elements [1], with all substances in their standard states under standard conditions [1].
 (b) $\text{Ca}(\text{s}) + \text{C}(\text{s, graphite}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{CaCO}_3(\text{s})$. Correct elements and coefficients [1]; correct states and exactly one mole of product [1].
 (c) $\frac{1}{2}\text{N}_2(\text{g}) + \frac{1}{2}\text{H}_2(\text{g}) \rightarrow \text{NH}_3(\text{g})$. Correct elements and coefficients [1]; correct states and exactly one mole of product [1].
 (d) 0 kJ mol^{-1} [1].

COMMON ERROR A formation equation starts from elements in their standard states, not from convenient compounds.

QUESTION 5 | TEST IT

7 marks

- (a) The enthalpy change when one mole of a substance burns completely in oxygen [1], with reactants and products in standard states [1] under standard conditions [1].

QUESTION 5 | TEST IT

7 marks

- (b) $\text{C}_3\text{H}_8\text{O(l)} + 4\text{O}_2\text{(g)} \rightarrow 3\text{CO}_2\text{(g)} + 3\text{H}_2\text{O(l)}$. Correct balancing [1]; correct standard states and complete-combustion products [1].
 (c) The definition is based on combustion of exactly one mole of propanone [1].
 (d) Formation of CO is incomplete combustion, so the defined complete-combustion products have not been formed [1].

QUESTION 6 | TEST IT

7 marks

- (a) $\Delta T = 26.75 - 18.35 = 8.40 \text{ K}$ [1].
 $q = mc\Delta T = 125.0 \times 4.18 \times 8.40$ [1] = 4389 J [1] = 4.39 kJ [1].
 The sign for the water is positive because it gains heat [1].
 (b) $m = (35.0 + 25.0) \times 1.00 = 60.0 \text{ g}$ [1]. Both solutions form the mixture whose temperature changes [1].

QUESTION 7 | TEST IT

6 marks

- (a) $q(\text{solution}) = 102.5 \times 4.18 \times 7.60 = +3256.22 \text{ J}$ [1] for the correct substitution and [1] for the value.
 $n(\text{LiCl}) = 2.50 \div 42.4 = 0.0589623 \text{ mol}$ [1].
 $q(\text{reaction}) = -3.25622 \text{ kJ}$ [1]; $\Delta H = -3.25622 \div 0.0589623 = -55.2 \text{ kJ mol}^{-1}$ [1].
 (b) The dissolving process releases heat to the solution, so it is exothermic and ΔH is negative [1].
COMMON ERROR The 102.5 g solution mass belongs in $q = mc\Delta T$. The 2.50 g LiCl mass belongs in $n = m/M$.

QUESTION 8 | TEST IT

7 marks

- (a) Maximum absolute uncertainty in $\Delta T = \pm(0.15 + 0.15) = \pm 0.30 \text{ K}$ [1].
 $\Delta T = 7.60 \text{ K}$, so percentage uncertainty = $0.30 \div 7.60 \times 100$ [1] = 3.9% [1].
 (b) Mass loss = $112.84 - 112.19 = 0.65 \text{ g}$ [1]. Maximum absolute uncertainty = $\pm 0.02 \text{ g}$ [1].
 Percentage uncertainty = $0.02 \div 0.65 \times 100 = 3.1\%$ [1].
 (c) $-64.3 \text{ kJ mol}^{-1}$ [1].

SECTION B EXAM STYLE ANSWERS

Model answers show the calculation stages and terminology expected by AQA.

QUESTION 9 | EXAM STYLE

9 marks

- (a) 100.0 g [1].
 (b) $\Delta T = 18.70 - 22.50 = -3.80 \text{ K}$ [1].
 $q(\text{solution}) = 100.0 \times 4.18 \times (-3.80)$ [1] = -1588.4 J = -1.5884 kJ [1].
 (c) $n(\text{NH}_4\text{NO}_3) = 4.00 \div 80.0 = 0.0500 \text{ mol}$ [1].
 $q(\text{dissolving}) = +1.5884 \text{ kJ}$ [1]; $\Delta H = 1.5884 \div 0.0500 = +31.8 \text{ kJ mol}^{-1}$ [1].
 (d) Heat from the room reduces the measured temperature fall and the magnitude of q [1], so the calculated positive ΔH is too small [1].

QUESTION 10 | EXAM STYLE

10 marks

- (a) $n(\text{HCl}) = 0.800 \times 35.0 \div 1000 = 0.0280 \text{ mol}$ [1].
 $n(\text{NaOH}) = 1.00 \times 25.0 \div 1000 = 0.0250 \text{ mol}$ [1]. NaOH is limiting in the 1:1 reaction [1].
 (b) $m = 60.0 \text{ g}$ [1]. $q(\text{solution}) = 60.0 \times 4.18 \times 5.80 = +1454.64 \text{ J} = +1.45464 \text{ kJ}$ [1].
 (c) $n(\text{H}_2\text{O}) = 0.0250 \text{ mol}$ [1]. The reaction releases the heat, so $q(\text{reaction})$ is negative [1].
 $\Delta H = -1.45464 \div 0.0250 = -58.2 \text{ kJ mol}^{-1}$ [1].
 (d) Any two valid assumptions, including: density is 1.00 g cm^{-3} ; the solution has the supplied c ; heat transfer to the cup and room is negligible; both reactants begin at the same temperature [2].

QUESTION 11 | EXAM STYLE

10 marks

- (a) $n(\text{H}_2\text{SO}_4) = 0.750 \times 30.0 \div 1000 = 0.0225 \text{ mol}$ [1].
 $n(\text{KOH}) = 0.600 \times 55.0 \div 1000 = 0.0330 \text{ mol}$ [1]. Comparing $n \div \text{coefficient}$ gives 0.0225 and 0.0165, so KOH limits [1]. $n(\text{H}_2\text{O}) = 0.0330 \text{ mol}$ [1].
 (b) $q(\text{solution}) = 85.0 \times 4.18 \times 4.20 = 1492.26 \text{ J}$ [1].
 $q(\text{reaction}) = -1.49226 \text{ kJ}$ [1]; $\Delta H = -1.49226 \div 0.0330 = -45.2 \text{ kJ mol}^{-1}$ [1].
 (c) Both solution volumes form the mixture whose temperature changes, so their masses are combined [1].
 (d) Some released heat warms the cup or escapes to the room [1], so the measured ΔT and q are too small and ΔH is less negative [1].

QUESTION 12 | EXAM STYLE

8 marks

- (a) $n(\text{CuSO}_4) = 0.400 \times 75.0 \div 1000 = 0.0300 \text{ mol}$ [1].
 $q(\text{solution}) = 75.0 \times 4.18 \times 19.8$ [1] = 6207.3 J = 6.2073 kJ [1].
 The reaction is exothermic [1], so $\Delta H = -6.2073 \div 0.0300 = -207 \text{ kJ mol}^{-1}$ [1].
 (b) The supplied heat-capacity model applies to the solution; excess undissolved iron does not have the same specific heat capacity [1].
 (c) The chosen m and calculated q would be too large [1], so the magnitude of ΔH would be too large and the result would be too negative [1].

QUESTION 13 | EXAM STYLE

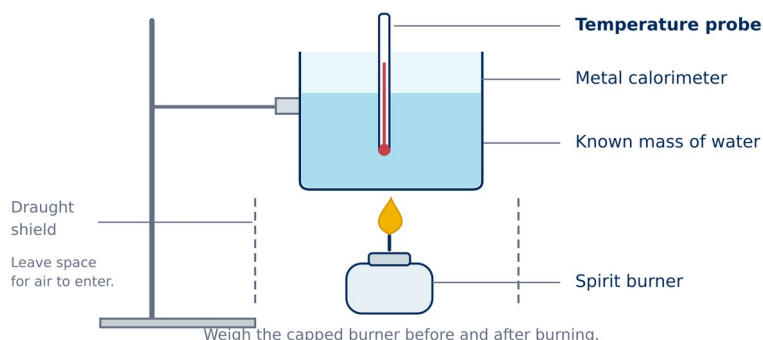
13 marks

(a) Fuel burned = $158.72 - 157.85 = 0.870 \text{ g}$ [1]. $\Delta T = 39.8 - 17.6 = 22.2 \text{ K}$ [1].

$q(\text{water}) = 200.0 \times 4.18 \times 22.2 \text{ [1]} = 18.5592 \text{ kJ}$ [1].

$n(\text{fuel}) = 0.870 \div 74.0 = 0.0117568 \text{ mol}$ [1].

$\Delta_c H = -18.5592 \div 0.0117568 = -1.58 \times 10^3 \text{ kJ mol}^{-1}$ [1].



(b) Percentage captured = $1578.6 \div 2676 \times 100 \text{ [1]} = 59.0\%$ [1].

(c) Heat transferred to the room does not warm the water, so ΔT and q are too small [1]. The experimental value therefore has too small a magnitude and is less negative [1].

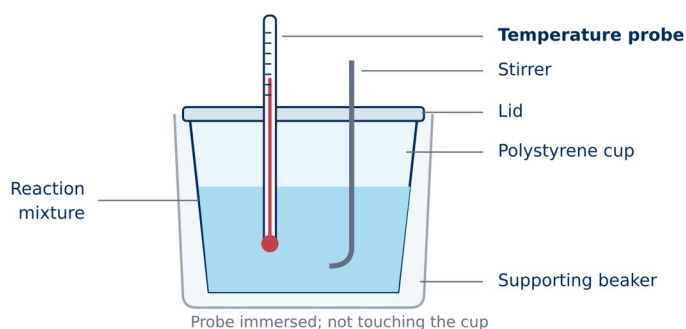
(d) Capping and prompt weighing reduce fuel evaporation [1], which would otherwise make the apparent mass burned too large [1].

(e) Oxygen must reach the flame to promote complete combustion [1].

QUESTION 14 | EXAM STYLE

10 marks

(a) The polystyrene cup reduces heat transfer [1]; the lid reduces heat transfer and evaporation [1]; gentle stirring makes the temperature more uniform [1].



(b) Measure known volumes and concentrations of the solutions [1]. Record temperature at fixed intervals before mixing [1]. Add the second reactant promptly, replace the lid and stir [1]. Continue readings after the maximum [1]. Plot temperature against time and extrapolate the later cooling line back to the mixing time [1].

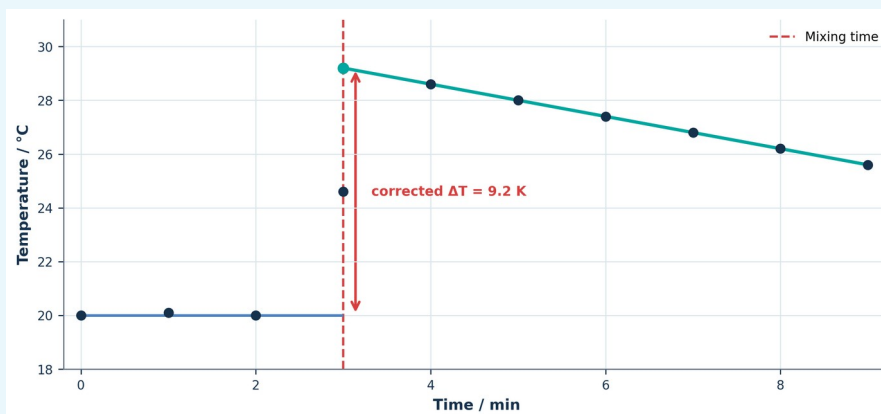
(c) Contact with the cup can give a temperature that is not representative of the well-mixed reaction mixture [1].

(d) Wear eye protection when handling the acid and alkali [1]. Accept another relevant risk-assessed precaution.

QUESTION 15 | EXAM STYLE

7 marks

(a) A straight cooling line is drawn and extended to 3.0 min [1]. The corrected maximum is 29.2°C [1], so corrected $\Delta T = 29.2 - 20.0 = 9.2 \text{ K}$ [1].



(b) Measured rise = $28.6 - 20.0 = 8.6 \text{ K}$ [1].

Percentage underestimate = $(9.2 - 8.6) \div 9.2 \times 100 = 6.5\%$ [1].

(c) Heat is already being lost before the measured maximum is recorded [1]. Extrapolation estimates the temperature the mixture would have reached at the mixing time without that loss [1].

QUESTION 16 | EXAM STYLE

8 marks

- (a) Trial 3 is anomalous [1] because it is far from the two concordant values in trials 1 and 2 [1].
 (b) Mean = $(-52.1 - 52.3) \div 2 = -52.2 \text{ kJ mol}^{-1}$ [1].
 (c) Absolute difference = $57.3 - 52.2 = 5.1 \text{ kJ mol}^{-1}$ [1].
 Percentage difference = $5.1 \div 57.3 \times 100 = 8.9\%$ [1].
 (d) Heat loss affects each run in the same direction [1], so the mean remains biased and less negative [1].
 (e) Use more insulation or a well-fitting lid [1].

SECTION C HARDER MIXED ANSWERS

These solutions make each choice, assumption and error direction explicit.

QUESTION 17 | HARDER MIXED

12 marks

- (a) $q(\text{solution}) = 100.0 \times 4.18 \times 2.95$ [1] = 1233.1 J [1].
 $q(\text{calorimeter}) = 42.0 \times 2.95 = 123.9 \text{ J}$ [1]. Total heat gain = 1357.0 J [1].
 (b) $n(\text{Na}_2\text{CO}_3) = 5.00 \div 106.0 = 0.0471698 \text{ mol}$ [1]. The dissolving process released -1.3570 kJ [1].
 $\Delta H = -1.3570 \div 0.0471698 = -28.8 \text{ kJ mol}^{-1}$ [1].
 (c) Using only $q(\text{solution})$, $\Delta H = -1.2331 \div 0.0471698$ [1] = -26.1 kJ mol⁻¹ [1].
 (d) Percentage understatement = $(28.768 - 26.142) \div 28.768 \times 100$ [1] = 9.13% [1].
 (e) The calorimeter absorbs some released energy, so omitting it underestimates the total heat released [1].

QUESTION 18 | HARDER MIXED

11 marks

- (a) $q(\text{solution}) = 100.0 \times 4.18 \times 6.20$ [1] = 2591.6 J = 2.5916 kJ [1].
 (b) $n(\text{H}_2\text{O}) = 2.5916 \div 57.1 = 0.045387 \text{ mol}$ [1] for the route and [1] for the amount.
 $n(\text{NaOH}) = 1.00 \times 50.0 \div 1000 = 0.0500 \text{ mol}$ [1]. Since less than 0.0500 mol reacts in the 1:1 equation, HCl is limiting [1].
 (c) $50.0 \text{ cm}^3 = 0.0500 \text{ dm}^3$ [1]. $c(\text{HCl}) = 0.045387 \div 0.0500$ [1] = 0.908 mol dm⁻³ [1].
 (d) Halving m halves q and gives 0.454 mol dm⁻³ [1]. The concentration is too low because the heat gain and inferred reacting amount are halved [1].

QUESTION 19 | HARDER MIXED

10 marks

- (a) $q(\text{mixture}) = 77.5 \times 4.18 \times (-1.40)$ [1] = -453.53 J [1].
 (b) $q(\text{dissolving}) = +0.45353 \text{ kJ}$ [1]. $n(\text{KCl}) = 0.45353 \div 17.2 = 0.026368 \text{ mol}$ [1].
 (c) Mass of pure KCl = $0.026368 \times 74.6 = 1.967 \text{ g}$ [1].
 Percentage purity = $1.967 \div 2.50 \times 100$ [1] = 78.7% [1].
 (d) Heat entering reduces the observed temperature fall and magnitude of q [1], so the inferred amount of KCl and percentage purity are too low [1].
 (e) The impurity is inert and causes no heat change of its own [1].

QUESTION 20 | HARDER MIXED

10 marks

- (a) Compare $n \div \text{coefficient}$: H_2SO_4 gives 0.0300 and KOH gives 0.0250 [1]. KOH is limiting [1].
 (b) $m = 20.0 + 50.0 = 70.0 \text{ g}$ [1]. $q(\text{solution}) = 70.0 \times 4.18 \times 6.00 = 1755.6 \text{ J} = 1.7556 \text{ kJ}$ [1].
 (c) The reaction is exothermic, so $q(\text{reaction})$ is negative [1]. The mole basis is $n(\text{H}_2\text{O}) = 0.0500 \text{ mol}$ [1].
 (d) $\Delta H = -1.7556 \div 0.0500$ [1] = -35.1 kJ mol⁻¹ [1].
 (e) Some released heat escapes or warms the apparatus [1], so the measured ΔT and magnitude of ΔH are too small [1].
COMMON ERROR Limiting reagent is determined from amount divided by coefficient, not by comparing the two raw mole values.

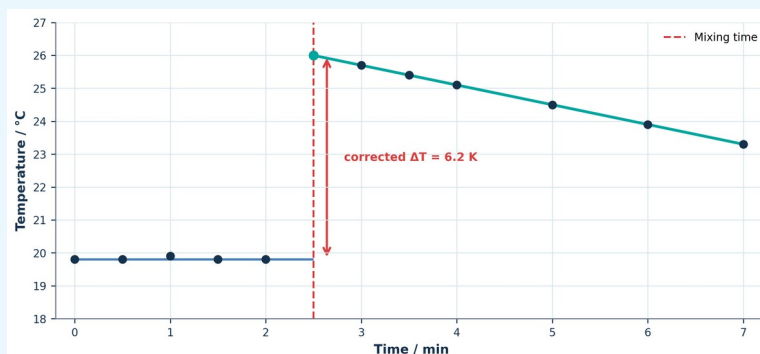
SECTION D STRETCH / SYNOPTIC ANSWERS

Full-credit solutions connect ideas without importing Energetics II content.

QUESTION 21 | STRETCH / SYNOPTIC

16 marks

- (a) The later straight cooling line is extrapolated to 2.5 min [1]. Initial temperature = 19.8 °C [1]; corrected maximum = 26.0 °C [1]; corrected $\Delta T = 6.2 \text{ K}$ [1].



- (b) $n(\text{HNO}_3) = 0.800 \times 40.0 \div 1000 = 0.0320 \text{ mol}$ [1]. $n[\text{Ba}(\text{OH})_2] = 1.00 \times 30.0 \div 1000 = 0.0300 \text{ mol}$ [1].

QUESTION 21 | STRETCH / SYNOPTIC**16 marks**

- Comparing $n \div \text{coefficient}$ gives 0.0160 and 0.0300, so HNO_3 limits [1]. $n(\text{H}_2\text{O}) = 0.0320 \text{ mol}$ [1].
- (c) $m = 70.0 \text{ g}$ [1]. $q(\text{solution}) = 70.0 \times 4.18 \times 6.2 = 1814.12 \text{ J}$ [1].
 $q(\text{reaction})$ is negative [1]; $\Delta H = -1.81412 \div 0.0320 = -56.7 \text{ kJ mol}^{-1}$ [1].
- (d) Maximum absolute uncertainty in $\Delta T = \pm 0.2 \text{ K}$ [1]. Percentage uncertainty = $0.2 \div 6.2 \times 100 = 3.2\%$ [1].
- (e) Uncorrected $\Delta T = 25.7 - 19.8 = 5.9 \text{ K}$, giving $\Delta H = -53.9 \text{ kJ mol}^{-1}$ [1]. It is less negative because heat loss lowered the measured maximum and q [1].

QUESTION 22 | STRETCH / SYNOPTIC**16 marks**

- (a) $q(\text{water}) = 150.0 \times 4.18 \times 12.0$ [1] = $7524 \text{ J} = 7.524 \text{ kJ}$ [1].
- (b) Energy released = $7.524 \div 0.550$ [1] = 13.68 kJ [1].
- (c) Methanol: $0.600 \div 32.0 \times 726 = 13.61 \text{ kJ}$ [1]. Ethanol: $0.600 \div 46.0 \times 1367 = 17.83 \text{ kJ}$ [1]. Propan-1-ol: $0.600 \div 60.0 \times 2021 = 20.21 \text{ kJ}$ [1]. The alcohol is methanol because 13.61 kJ matches 13.68 kJ [1].
- (d) Captured heat = $0.750 \times 13.6125 = 10.209 \text{ kJ}$ [1].
 $\Delta T = q \div (mc)$ [1] = $10\,209 \div (150.0 \times 4.18) = 16.3 \text{ K}$ [1].
- (e) Evaporation makes the apparent mass loss larger than the mass actually burned [1]. The calculated energy released per gram would therefore be too small [1], distorting the comparison with the candidate values [1].
- (f) Any two valid reasons, including: heat is transferred to the apparatus or room; combustion may be incomplete; fuel may evaporate; the experiment may not reproduce standard-state conditions [2].
- COMMON ERROR** The 55.0% refers to the fraction captured. Divide the measured heat by 0.550 to recover the estimated total released energy.

FINAL SPECIFICATION COVERAGE AUDIT

AQA / notes strand	Question coverage
3.1.4.1 enthalpy change	Q1-Q5, Q16: exothermic/endothermic change, signs, constant pressure, standard conditions, formation and combustion definitions and equations.
3.1.4.2 calculations	Q6-Q13, Q17-Q22: $q = mc\Delta T$, molar enthalpy, rearrangements, limiting reagents, solution, displacement, neutralisation and combustion.
Required Practical 2	Q8-Q16, Q19-Q22: apparatus, method, temperature-time correction, uncertainty, significant figures, repeats, limitations and improvements.
Boundary	Only a short supporting enthalpy-profile link in Q2. No Hess cycles or mean bond enthalpy calculations; these remain in Energetics II.