3.1 Energetics

Name: _____ Class: _____ Date: _____ **Total: 20 marks**

Objective

Build the skills to answer exam questions on **energetics** : exothermic and endothermic reactions, energy profiles, calorimetry with $Q = mc\Delta T$, molar enthalpy changes and bond energies.

You must be able to:

- tell exothermic and endothermic reactions apart
- calculate heat energy with $Q = mc\Delta T$
- calculate a molar enthalpy change ΔH in kJ/mol
- calculate ΔH from bond energies

1 Key knowledge and worked examples

Heat, moles and bonds

idea	rule or formula
exothermic	gives out heat; temperature rises; ΔH is negative
endothermic	takes in heat; temperature falls; ΔH is positive
heat energy	$Q = mc\Delta T$: m = mass of **water** (g), $c = 4.2$ J/g/°C, ΔT = temperature change
molar enthalpy change	$\Delta H = -Q \div$ moles (change J to kJ: $\div 1000$)
bond energies	ΔH = energy to break bonds − energy released making bonds

Worked example 1: heat energy. 50 g of water warms by 8.0 °C. Find Q.

1. Known: $m = 50$ g; $c = 4.2$; $\Delta T = 8.0$.
2. Rule: $Q = mc\Delta T$.
3. Numbers: $50 \times 4.2 \times 8.0$.
4. Result: 1680 **J** (1.68 kJ).

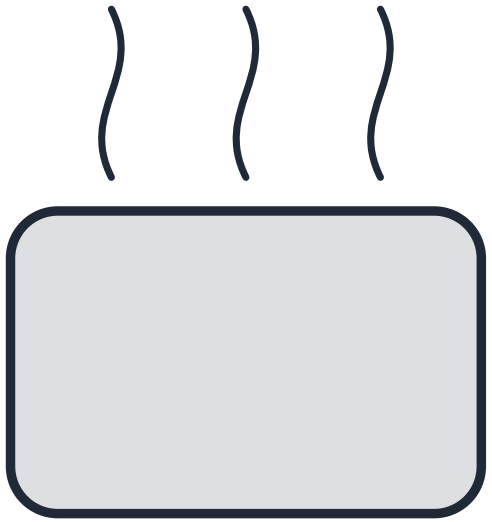
Worked example 2: bond energies. $H_2 + Cl_2 \rightarrow 2HCl$. Bond energies (kJ/mol): H−H 436, Cl−Cl 242, H−Cl 431.

1. Known: break 1 H−H and 1 Cl−Cl; make 2 H−Cl.
2. Rule: broken − made.
3. Numbers: $(436 + 242) - (2 \times 431) = 678 - 862$.
4. Result: $\Delta H = -184$ **kJ/mol** (exothermic).

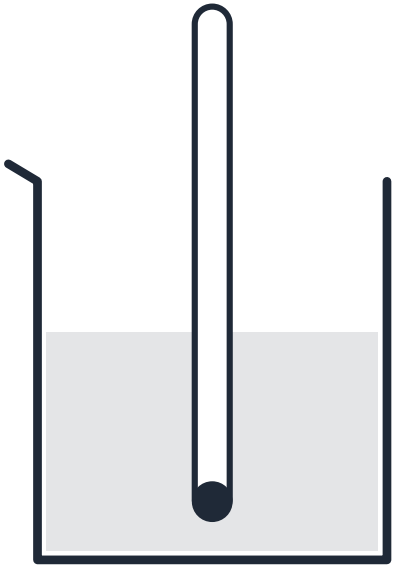
Answer shapes: give ΔH a sign and the unit kJ/mol.

2 Practice

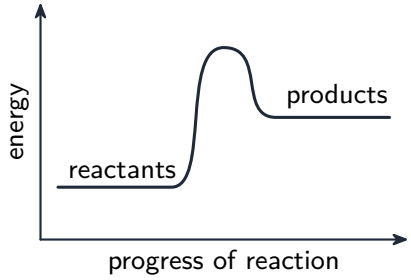
2.1 A hand warmer gets hot when it is opened. State whether the reaction inside is exothermic or endothermic, and what happens to the temperature of the surroundings. [2]



2.2 100 g of water warms from 20.0 °C to 26.0 °C. Calculate the heat energy taken in by the water ($c = 4.2$ J/g/°C). [2]

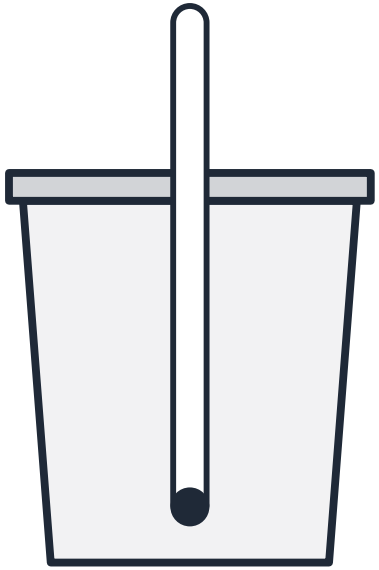


2.3 The energy profile for a reaction is shown. State whether the reaction is exothermic or endothermic, and give a reason. [2]



3 Exam-style questions

3.1 25 cm^3 of 1.0 mol/dm^3 hydrochloric acid is mixed with 25 cm^3 of 1.0 mol/dm^3 sodium hydroxide in a polystyrene cup. The temperature rises by 6.8 °C. Take the mass of the solution as 50 g. Calculate ΔH for the reaction in kJ/mol. [3]



3.2 A student burns ethanol in a spirit burner under a copper can holding 200 g of water. The results are shown. (M_r of ethanol = 46)



	value
mass of water	200 g
temperature rise	30 °C
mass of ethanol burned	1.15 g

(a) Calculate the heat energy taken in by the water. [2]

(b) Calculate the enthalpy change of combustion of ethanol in kJ/mol. [3]

(c) The value in a data book is −1367 kJ/mol. Explain why the student's value is smaller, and describe changes to the method that would give a more accurate value. [6]

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Solutions

Marking: **M1** a correct method, **A1** the correct answer, **B1** an independent mark.

Each answer shows the steps, then the **result**.

2.1

- exothermic
- the temperature of the surroundings rises

2.2

- $100 \times 4.2 \times 6.0$
- **2520** J

2.3

- endothermic
- the products have more energy than the reactants

3.1

- $Q = 50 \times 4.2 \times 6.8 = 1428$ J
- moles $= 1.0 \times 25 \div 1000 = 0.025$
- $\Delta H = -1.428 \div 0.025 = \mathbf{-57}$ kJ/mol

3.2

(a) $200 \times 4.2 \times 30 = \mathbf{25\,200}$ J

(b) moles ethanol $= 1.15 \div 46 = 0.025$; $25.2 \div 0.025 = 1008$; $\Delta H = \mathbf{-1008}$ kJ/mol

(c) points to include:

- heat is lost to the air around the can and the flame (An)
- heat is also used to warm the can itself (An)
- incomplete combustion: some ethanol burns to carbon monoxide or soot, giving out less energy (K)
- some ethanol evaporates from the wick, so the mass loss is too large (An)
- improvements: use a draught shield and a lid on the can; put the flame close to the can (App)
- stir the water; weigh the burner with its cap on, straight after putting the flame out (App)

Levels: **5–6** a full, logical answer with correct chemistry and correct terms; **3–4** several correct points, partly linked; **1–2** one or two relevant points.