Cambridge IGCSE · Chemistry 0620 · Notes

Chemical energetics

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Read first: **Formulae and equations**

What you need to know

By the end of this topic you should be able to:

1. **State what an exothermic reaction is**: it transfers **thermal energy to the surroundings**, so the temperature of the surroundings **goes up**.
2. **State what an endothermic reaction is**: it takes in **thermal energy from the surroundings**, so the temperature of the surroundings **goes down**.
3. **Read reaction pathway diagrams** for exothermic and endothermic reactions: say which kind of reaction a diagram shows, and why.
4. **Extended only:** state that the transfer of thermal energy during a reaction is called the **enthalpy change**, ΔH, of the reaction, and that ΔH is **negative** for exothermic reactions and **positive** for endothermic reactions.
5. **Extended only:** define the **activation energy**, E_a: the minimum energy that colliding particles must have in order to react.
6. **Extended only:** draw and label a reaction pathway diagram for an exothermic or endothermic reaction from information you are given, showing the reactants, the products, the enthalpy change ΔH and the activation energy E_a.
7. **Extended only:** state that **breaking bonds is endothermic** and **making bonds is exothermic**, and use this to explain why a reaction's enthalpy change is negative or positive.
8. **Extended only:** calculate the enthalpy change of a reaction from **bond energies**.

Understand it

Energy in and energy out

Every chemical reaction moves energy between the chemicals and their **surroundings** (the water, the test-tube, the air, your hand). You can't see the energy, but you can measure what it does: it changes the temperature of the surroundings.

- In an **exothermic** reaction the chemicals give out thermal (heat) energy. The surroundings warm up, so a thermometer in the mixture **rises**. Burning fuels, respiration, neutralising an acid with an alkali, and adding water to anhydrous copper(II) sulfate are all exothermic.
- In an **endothermic** reaction the chemicals take in thermal energy. The surroundings lose that energy and cool down, so the thermometer **falls**. Thermal decomposition (for example heating calcium carbonate), photosynthesis, and dissolving ammonium nitrate or ammonium chloride in water are endothermic.

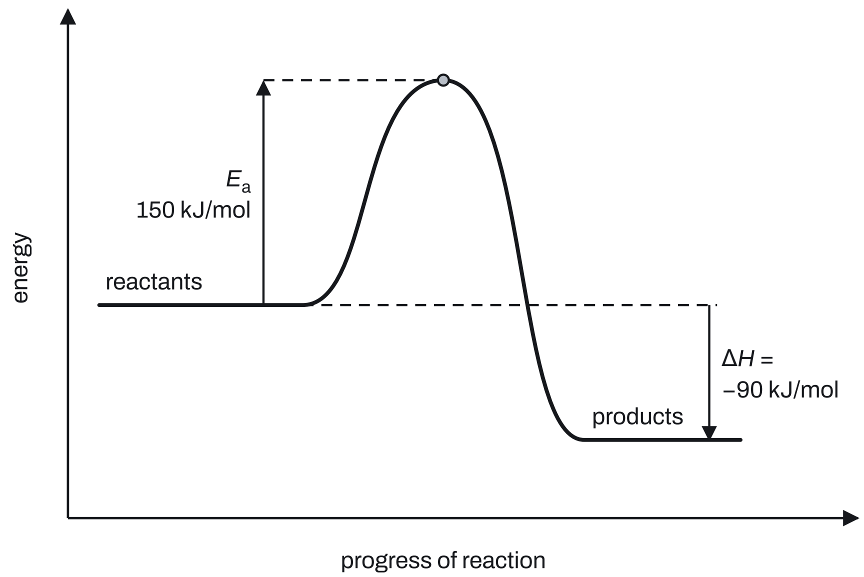
The thermometer measures the **surroundings**, not "the energy of the reaction". A falling temperature does not mean the reaction is losing energy; it means the reaction is **taking** energy from the water around it.

Comparing reactions. The **size of the temperature change** tells you how much energy moved. If one reaction goes from $20.0\ ^\circ\mathrm{C}$ to $27.5\ ^\circ\mathrm{C}$ (a rise of $7.5\ ^\circ\mathrm{C}$) and another goes from $20.0\ ^\circ\mathrm{C}$ to $16.0\ ^\circ\mathrm{C}$ (a fall of $4.0\ ^\circ\mathrm{C}$), the first is exothermic and the second is endothermic. With the same amounts and the same apparatus, a bigger rise means a more exothermic reaction and a bigger fall means a more endothermic one. Always work out the **change** (final minus initial); the final temperature on its own tells you nothing if the starting temperatures differ.

The same idea compares **fuels**. Burn each fuel under the same volume of water in a copper can and measure the temperature rise. More energy released means a bigger rise. If different masses of fuel were burned, divide the rise by the mass to compare them fairly per gram.

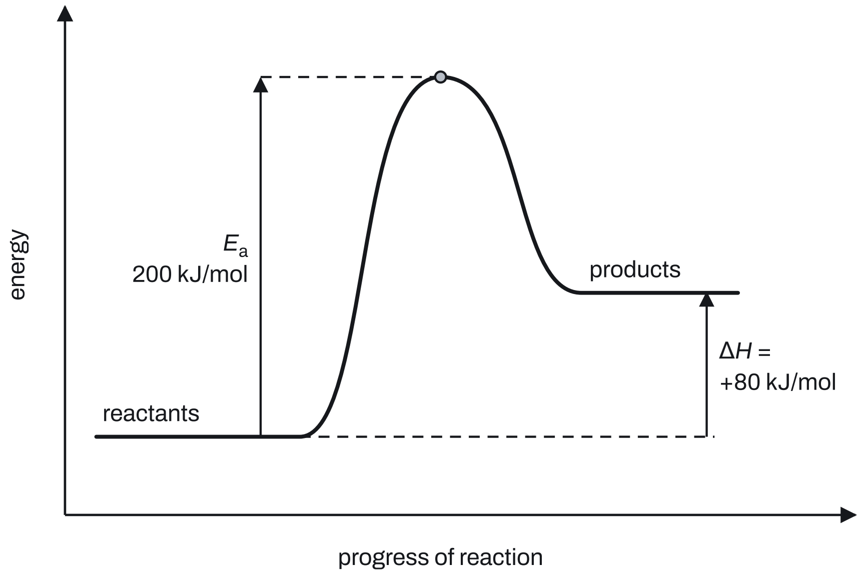
Reaction pathway diagrams

A **reaction pathway diagram** (also called an energy level diagram) shows the energy stored in the reactants and in the products. The vertical axis is **energy**; the horizontal axis is **progress of reaction**. The reactants are a short horizontal line on the left and the products a short horizontal line on the right.



- **Exothermic:** the products are **lower** than the reactants. The chemicals have lost energy, and that energy went into the surroundings as heat.

- **Endothermic:** the products are **higher** than the reactants. The chemicals have gained energy, taken from the surroundings.



When a question asks how a diagram shows the type of reaction, the answer is always about the **two levels**: "the products are at a lower energy than the reactants" (exothermic) or "higher" (endothermic). A definition of exothermic is not an answer to "how does the diagram show it".

Enthalpy change, ΔH (Extended)

Extended only: the thermal energy transferred during a reaction is called the **enthalpy change** of the reaction, ΔH, measured in kJ/mol. It is measured **from the chemicals' point of view**, so on the diagram the ΔH arrow always starts at the **reactants** and ends at the **products**.

- Exothermic: the chemicals lose energy, so ΔH is **negative**, for example $\Delta H = -90$ kJ/mol.
- Endothermic: the chemicals gain energy, so ΔH is **positive**, for example $\Delta H = +80$ kJ/mol.

If a reaction runs backwards, the arrow runs the other way: the reverse reaction has a ΔH of the **same size and opposite sign**. If $\mathrm{A} \rightarrow \mathrm{B}$ has $\Delta H = -90$ kJ/mol, then $\mathrm{B} \rightarrow \mathrm{A}$ has $\Delta H = +90$ kJ/mol. So if adding water to an anhydrous salt is exothermic, heating the hydrated salt to drive the water off is endothermic.

Activation energy (Extended)

Extended only: particles react only when they collide, and even then only if they hit hard enough to start breaking bonds. The **activation energy**, E_a, is the **minimum energy that colliding particles must have to react**. On the diagram it is the height of the hump, measured **from the reactants up to the peak**.

Every reaction has an activation energy, exothermic ones included: a fuel doesn't burn until you give it a spark or flame. The size of E_a is not linked to whether a reaction is exothermic or endothermic, and a reaction with a small E_a tends to be fast, because more collisions have enough energy. A catalyst lowers E_a; raising the temperature does **not** change E_a (it gives more particles enough energy to get over it).

Bond breaking and bond making (Extended)

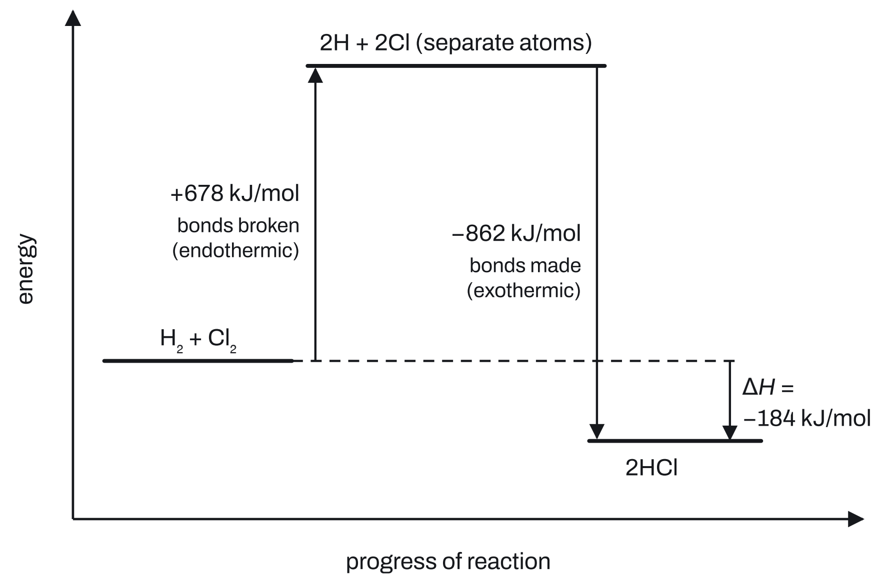
Extended only: in a reaction the bonds in the reactants break and new bonds form to make the products.

- **Breaking a bond takes energy in**, so bond breaking is **endothermic**. (Think of pulling two magnets apart.)
- **Making a bond gives energy out**, so bond making is **exothermic**.

The **bond energy** is the energy needed to break one mole of a particular bond, in kJ/mol; the same amount is released when one mole of that bond forms. The overall enthalpy change is the balance of the two:

$$\Delta H = (\text{energy taken in breaking bonds}) - (\text{energy given out making bonds})$$

- More energy given out making bonds than taken in breaking them: ΔH is negative, **exothermic**.
- More energy taken in breaking bonds than given out making them: ΔH is positive, **endothermic**.



For $H_2 + Cl_2 \rightarrow 2HCl$, with bond energies H–H 436, Cl–Cl 242 and H–Cl 431 kJ/mol:

- bonds broken: one H–H and one Cl–Cl, $436 + 242 = 678$ kJ taken in;
- bonds made: two H–Cl, $2 \times 431 = 862$ kJ given out;
- $\Delta H = 678 - 862 = -184$ kJ/mol. More energy is given out than taken in, so the reaction is exothermic.

Key facts and formulas

Nothing for this topic is printed in the exam. Bond energies are always given in the question; everything below is **Learn it**.

Formula	Status
Exothermic: thermal energy transferred **to** the surroundings; temperature of the surroundings **increases**	Learn it

Formula	Status
Endothermic: thermal energy taken in **from** the surroundings; temperature of the surroundings **decreases**	Learn it
Pathway diagram: products **below** reactants = exothermic; products **above** reactants = endothermic	Learn it
Temperature change $=$ final temperature $-$ initial temperature; bigger change = more energy transferred	Learn it
Comparing fuels: energy per gram compared by $\frac{\text{temperature rise}}{\text{mass of fuel burned}}$	Learn it
Extended only: ΔH = enthalpy change = thermal energy transferred during a reaction; negative for exothermic, positive for endothermic	Learn it
Extended only: reverse reaction: same size of ΔH, opposite sign	Learn it
Extended only: E_a = minimum energy that colliding particles must have to react; arrow from reactants **up to the peak**	Learn it
Extended only: ΔH arrow from the reactants level **to** the products level (down for exothermic, up for endothermic)	Learn it
Extended only: bond breaking is endothermic; bond making is exothermic	Learn it
Extended only: $\Delta H = \text{bonds broken} - \text{bonds made}$ (each bond energy $\times$ number of those bonds)	Learn it

How to do it

In the questions we have tagged for this topic (132 questions, 2021–2025, on Papers 1–4), the types came up this often. 24 of the 132 questions mix types, so the counts add up to more than 132.

Question type	Questions
Interpret a reaction pathway diagram	38
Link exothermic or endothermic to temperature change, energy transfer and the sign of ΔH	29
Calculate ΔH from bond energies	25
State a definition (exothermic, endothermic, ΔH)	17
Complete or draw a reaction pathway diagram	15
Identify, label or define activation energy and ΔH	12
Classify a named reaction as exothermic or endothermic	10
Calculate an unknown bond energy from ΔH	7

Question type	Questions
Explain exothermic or endothermic using bond breaking and bond making	7
Compare fuels from the temperature rise	5

1. Interpret a reaction pathway diagram (38 questions)

1. **Find the reactants line and the products line.** They are labelled with words or formulae.
2. **Compare their heights.** Products lower: **exothermic**, heat given out, surroundings get warmer. Products higher: **endothermic**, heat taken in, surroundings get cooler.
3. **Give the reason from the diagram**: "the energy of the products is lower (higher) than the energy of the reactants". This is the second mark in a written answer.

Variations you will meet:

- **Multiple choice pairs** the type with "thermal energy given out to (taken in from) the surroundings" or with a temperature rise or fall. Decide the type first, then pick the matching pair.
- **"Which diagram shows the most energy released"** (or the largest temperature rise) when the scales are the same: pick the exothermic diagram with the **biggest drop** from reactants to products. The height of the hump doesn't matter.
- **Match an equation to its diagram.** Combustion is exothermic, so choose the diagram with the products (for example $\mathrm{CO_2 + H_2O}$) lower.
- **The reverse reaction.** If forming a hydrated salt from the anhydrous salt and water gives out energy, the reverse (hydrated salt $\rightarrow$ anhydrous salt + water) takes energy in: its diagram has the anhydrous salt + water **higher**, on the right.

2. Link exothermic or endothermic to temperature change, energy transfer and the sign of ΔH (29 questions)

1. For each experiment work out **change = final – initial**.
2. **A rise (positive change) means exothermic; a fall (negative change) means endothermic.**
3. Rank by the **size** of the change:
   - "greatest temperature change": the biggest number, ignoring whether it is a rise or fall;
   - "most exothermic": the biggest **rise**; "least exothermic": the smallest rise;
   - "most endothermic": the biggest **fall**.

For example, experiments with $22 \rightarrow 30$, $21 \rightarrow 15$, $20 \rightarrow 31$ and $23 \rightarrow 13$ (°C) have changes $+8$, -6, $+11$ and -10. The greatest change and the most exothermic is the third ($11\ ^\circ\mathrm{C}$); the most endothermic is the fourth; the least exothermic is the first.

Variations you will meet:

- **Dissolving a solid**: a fall in temperature (ammonium nitrate, ammonium chloride) means the change is endothermic, energy is **absorbed** from the water.

- **Statements** such as "in an endothermic reaction the temperature of the surroundings increases": false. Check each statement against the two definitions.
- **Extended only: the sign of ΔH.** Exothermic $\leftrightarrow$ ΔH negative $\leftrightarrow$ surroundings get warmer. Endothermic $\leftrightarrow$ ΔH positive $\leftrightarrow$ surroundings get cooler.

3. Calculate ΔH from bond energies (25 questions)

Extended only.

1. **Draw or read the displayed formulae** and **count every bond** in the reactants and in the products, multiplying by the numbers in front of each formula. A double bond (C=C, O=O, C=O) is one bond with its own bond energy, not two single bonds.
2. **Energy taken in (bonds broken)** = sum of (bond energy × number of those bonds) in the **reactants**.
3. **Energy given out (bonds made)** = the same sum for the **products**.
4. $\Delta H = \text{broken} - \text{made}$. Give the **sign**: negative is exothermic, positive is endothermic.

Paper 4 usually sets these three steps as three separate answer lines, a mark each. Write the two totals as **positive** numbers, then subtract.

Variations you will meet:

- **"Per mole of fuel"** when the equation has two moles of it: work out ΔH for the equation as written, then divide by 2.
- **Choose the right expression** (multiple choice): it is (bonds in reactants) $-$ (bonds in products), not the other way round.
- **Bonds formed from atoms only**, for example forming 1 mol of $\mathrm{NH_3}$ from separate N and H atoms: nothing is broken, three N–H bonds form, so $\Delta H = -3 \times 391 = -1173$ kJ.
- **Extra bonds in the table** that the reaction doesn't use (such as C–O when no C–O bond appears): ignore them.
- **Follow-up:** "Is the reaction exothermic or endothermic? Give a reason." Use the sign, or say which total is bigger (see type 9).

4. State a definition (17 questions)

Learn these, word for word in meaning:

- **Exothermic**: a reaction that transfers **thermal energy to the surroundings** (the surroundings get warmer).
- **Endothermic**: a reaction that takes in **thermal energy from the surroundings** (the surroundings get cooler).
- **Extended only:** ΔH is the **enthalpy change** of the reaction: the thermal energy transferred during the reaction. A **negative** value shows the forward reaction is exothermic.
- **Extended only:** the reverse reaction's ΔH has the same size and the opposite sign (for example $-105 \rightarrow +105$ kJ/mol). Write the $+$ sign.

"Energy is released" alone is not enough for exothermic: say **thermal** (or heat) energy, and say where it goes (the surroundings). Some Core questions give the sentence with a gap to fill: the missing word is usually "surroundings".

5. Complete or draw a reaction pathway diagram (15 questions)

1. **Label the vertical axis "energy"** if it is blank.
2. **Put the reactants on the left line and the products on the right line**, written on or just above the line. Use the formulae from the equation if the question gives them.
3. **Extended only: place the products line** below the reactants for an exothermic reaction (negative ΔH) and above for an endothermic one. Draw a smooth hump between them, higher than both lines.
4. **Extended only: activation energy arrow**: vertical, from the **reactants level up to the top of the hump**, pointing **up**, labelled E_a.
5. **Extended only: ΔH arrow**: vertical, from the **reactants level to the products level**, pointing **down** (exothermic) or **up** (endothermic), labelled ΔH.

Put exactly one arrowhead on each arrow and don't add extra arrows or labels: examiners reject arrows pointing both ways and formulae written on the arrows.

6. Identify, label or define activation energy and ΔH (12 questions)

Extended only.

1. E_a is the arrow or gap from the **reactants** up to the **peak**. It is not the gap from the products to the peak, and not the height of the peak above the axis.
2. ΔH is the arrow or gap from the **reactants** to the **products**.
3. **Definition of E_a**: the minimum energy that colliding particles must have to react. Include "minimum" and "colliding particles".

Variations you will meet:

- **ΔH from a labelled diagram**: if the diagram gives $E_a = +180$ kJ/mol (reactants up to the peak) and the energy released from the peak down to the products as -300 kJ/mol, then $\Delta H = 180 - 300 = -120$ kJ/mol.
- **Statements about E_a**: particles must collide with at least E_a to react (true). E_a is not always larger for exothermic reactions, and it does not depend on whether ΔH is positive or negative.
- **"One way to change the activation energy"**: add a **catalyst**, which lowers it. Heating does not change E_a.

7. Classify a named reaction as exothermic or endothermic (10 questions)

Learn these examples:

Exothermic (thermal energy given out)	Endothermic (thermal energy taken in)
Burning any fuel (combustion), including hydrogen, methane and a match	Thermal decomposition, for example $CaCO_3 \rightarrow CaO + CO_2$
Metals burning in oxygen	Photosynthesis
Neutralisation (acid + alkali)	Dissolving ammonium nitrate or ammonium chloride
Respiration	Heating a hydrated salt to remove its water
Adding water to anhydrous copper(II) sulfate	Breaking a bond, for example $Cl_2 \rightarrow 2Cl$
Forming a bond, for example $H + H \rightarrow H_2$	

Burning a fuel **always** releases thermal energy, whatever its products. A reaction that needs **continuous heating** to keep going (like decomposing calcium carbonate) is endothermic, and it is still a chemical change.

8. Calculate an unknown bond energy from ΔH (7 questions)

Extended only.

1. **Add up every bond energy you know** on the side without the unknown.
2. **Use** $\Delta H = \text{broken} - \text{made}$ to find the total on the other side:
   - unknown in the products: $\text{made} = \text{broken} - \Delta H$ (with $\Delta H = -54$, this means **adding** 54);
   - unknown in the reactants: $\text{broken} = \Delta H + \text{made}$.
3. **Take away the known bonds** on that side, leaving only the unknown bonds.
4. **Divide by how many** of the unknown bond there are.

"230 kJ of energy is released" means $\Delta H = -230$ kJ/mol. Keep the sign in the equation and the rearranging looks after itself.

9. Explain exothermic or endothermic using bond breaking and bond making (7 questions)

Extended only.

- Bond **breaking** takes energy **in** (endothermic); bond **making** gives energy **out** (exothermic).
- **Exothermic**: "more energy is **released** when the bonds in the products are **formed** than is **needed** to **break** the bonds in the reactants."
- **Endothermic**: "more energy is **needed** to break the bonds in the reactants than is **released** when the bonds in the products are formed." Photosynthesis is an example.
- Never write that forming bonds "needs" or "uses" energy.

Multiple-choice statements mix this with the diagram and the temperature; check each part separately.

10. Compare fuels from the temperature rise (5 questions)

1. **Work out each temperature rise** = final – initial. Don't compare final temperatures if the starting temperatures differ.
2. If the **same mass** of each fuel was burned (or the same time), the biggest rise released the most energy.
3. If **different masses** were burned, work out $\frac{\text{rise}}{\text{mass}}$ for each: the biggest value releases the most energy **per gram**, the smallest the least.

Worked examples

The bold codes show where marks are earned, as in Cambridge mark schemes. The number is how many marks.

- **M1** method mark: for a correct method, even if the numbers slip.
- **A1** accuracy mark: for a correct answer; it needs the M mark before it.
- **B1** mark for a correct result on its own, with no method needed.

Example 1

A student adds a small amount of four different substances, A, B, C and D, to separate $25\ \text{cm}^3$ samples of dilute hydrochloric acid and records the temperatures.

substance	initial temperature / °C	final temperature / °C
A	19.5	26.0
B	20.0	14.5
C	21.0	23.5
D	20.5	12.0

(a) State which substances react endothermically. **[1]**

(b) State which experiment shows the greatest temperature change. **[1]**

(c) State which experiment is the most exothermic. **[1]**

(d) A reaction pathway diagram with two horizontal lines is drawn for the reaction of substance D. State how you would label it. **[2]**

(a) Changes: A $+6.5$, B -5.5, C $+2.5$, D -8.5 °C. The temperature falls for **B and D**, so they are endothermic. **B1**

(b) The largest change, ignoring the sign, is $8.5\ ^\circ\text{C}$: **D**. **B1**

(c) The largest **rise** is $6.5\ ^\circ\mathrm{C}$: **A**. **B1** (D has the biggest change, but it is a fall.)

(d) The vertical axis is labelled **energy**. **B1** Reaction D is endothermic, so the **reactants** go on the **lower** line on the left and the **products** on the **higher** line on the right. **B1**

Example 2

Sulfur dioxide reacts with oxygen to form sulfur trioxide.

$$2\mathrm{SO_2(g)} + \mathrm{O_2(g)} \rightarrow 2\mathrm{SO_3(g)} \qquad \Delta H = -196\ \mathrm{kJ/mol}$$

The activation energy of this reaction is 250 kJ/mol.

(a) State how the value of ΔH shows that the reaction is exothermic. **[1]**

(b) Draw a reaction pathway diagram for the reaction. Show the reactants and products, an arrow labelled E_a for the activation energy and an arrow labelled ΔH for the enthalpy change. **[4]**

(c) Define the term activation energy. **[1]**

(d) Deduce the enthalpy change for the reverse reaction, $2\mathrm{SO_3} \rightarrow 2\mathrm{SO_2} + \mathrm{O_2}$. **[1]**

(a) ΔH is **negative**. **B1**

(b) Label the vertical axis energy and draw:

- a line on the left labelled $2\mathrm{SO_2} + \mathrm{O_2}$ and a **lower** line on the right labelled $2\mathrm{SO_3}$, joined by a hump; **M1** for the products below the reactants, **M1** for both sets of formulae on the right lines;
- an upward arrow from the $2\mathrm{SO_2} + \mathrm{O_2}$ level to the top of the hump, labelled E_a (the hump rises 250 kJ/mol above the reactants); **M1**
- a downward arrow from the reactants level to the products level, labelled ΔH $(196$ kJ/mol). **M1**

Seen from the products, the peak is $250 + 196 = 446$ kJ/mol above them: the hump must stand well above both lines.

(c) The minimum energy that colliding particles must have in order to react. **B1**

(d) Same size, opposite sign: $\Delta H = +196$ kJ/mol. **B1** (the $+$ sign is needed).

Example 3

Methanol, $\mathrm{CH_3OH}$, burns in oxygen.

$$2\mathrm{CH_3OH} + 3\mathrm{O_2} \rightarrow 2\mathrm{CO_2} + 4\mathrm{H_2O}$$

bond	C–H	C–O	O–H	O=O	C=O
bond energy in kJ/mol	410	360	460	496	805

(a) Use the bond energies to calculate the enthalpy change for the equation as written. Use these steps: energy needed to break the bonds; energy released when the bonds form; the enthalpy change. **[3]**

(b) Calculate the enthalpy change when 1 mol of methanol burns. **[1]**

(c) Explain, in terms of bond breaking and bond making, why the reaction is exothermic. **[2]**

(a) Displayed formulae: methanol has 3 C–H, 1 C–O and 1 O–H; oxygen O=O; carbon dioxide O=C=O (2 C=O); water H–O–H (2 O–H).

Bonds broken: $2 \times (3 \times 410 + 360 + 460) + 3 \times 496 = 2 \times 2050 + 1488 = 5588$ kJ. **B1**

Bonds made: $2 \times (2 \times 805) + 4 \times (2 \times 460) = 3220 + 3680 = 6900$ kJ. **B1**

$\Delta H = 5588 - 6900 = -1312$ kJ/mol. **B1** (with the minus sign).

(b) The equation has 2 mol of methanol, so per mole $-1312 \div 2 = -656$ kJ/mol. **B1**

(c) More energy is released when the bonds in the products are formed (6900 kJ) **B1** than is needed to break the bonds in the reactants (5588 kJ). **B1**

Example 4

Nitrogen and hydrogen react to make ammonia.

$$\mathrm{N \equiv N + 3\,H\text{–}H \rightarrow 2\,NH_3} \qquad \Delta H = -92\ \mathrm{kJ/mol}$$

The bond energy of H–H is 436 kJ/mol and of N–H is 391 kJ/mol. Each $\mathrm{NH_3}$ molecule has three N–H bonds.

Calculate the bond energy of the $\mathrm{N \equiv N}$ bond. Use these steps: energy released forming the bonds in the products; energy needed to break the bonds in the reactants; the bond energy of N≡N. **[3]**

Bonds made: $2 \times 3 \times 391 = 2346$ kJ. **B1**

Bonds broken $= \Delta H + \mathrm{made} = -92 + 2346 = 2254$ kJ. **M1**

$\mathrm{N \equiv N} = 2254 - 3 \times 436 = 2254 - 1308 = 946$ kJ/mol. **A1**

Check: broken $= 946 + 1308 = 2254$, and $2254 - 2346 = -92$. ✓

Common mistakes

- **Mixing up exothermic and endothermic.** Examiners see this every session, on every paper. Tie each word to the thermometer: **ex**othermic, heat **ex**its to the surroundings, temperature **up**; endothermic, heat goes **in**, temperature **down**.
- **Reading a temperature rise as endothermic.** "The temperature of the mixture increases" means heat was given out: exothermic.
- **Comparing final temperatures instead of changes.** A mixture that ends at a higher temperature has not necessarily changed the most. Always subtract.

- **Defining instead of using the diagram.** "How does the diagram show the reaction is exothermic?" wants "the products have less energy than the reactants", not "it gives out heat".
- **Vague definitions.** "Releases energy" misses the mark: say **thermal** (heat) energy transferred **to the surroundings**. For ΔH, write "enthalpy change of the reaction", not just "energy" or "enthalpy".
- **Diagram slips**: reactants and products swapped, labels written away from the lines, formulae written on the arrows, arrows pointing both ways, or extra arrows. Examiners reject all of these.
- **Measuring E_a from the wrong place.** It starts at the **reactants** level, not the products and not the axis. Confusing E_a with ΔH is a common wrong option.
- **Believing heat or temperature changes E_a.** Only a catalyst changes it.
- **Subtracting the wrong way round.** $\Delta H = \text{broken} - \text{made}$. "Made – broken" gives the right number with the wrong sign, a common wrong answer.
- **Missing or miscounted bonds.** Count every bond in every molecule, times the number in front of the formula: $4\text{H}_2\text{O}$ has 8 O–H bonds. Remember the O–H in an alcohol.
- **Unknown bond energy steps**: subtracting the size of ΔH when it should be added, forgetting to take away the known bonds, and forgetting to divide by the number of unknown bonds.
- **Saying bond forming needs energy.** Forming bonds **releases** energy; only breaking bonds needs it.
- **Dropping the sign.** A ΔH without $+$ or $-$ usually loses the mark, and the reverse reaction's value must show $+$.

How to get the marks

- **"State"** (for example "State the meaning of exothermic"): one precise sentence with the key words: thermal energy, transferred, surroundings.
- **"Explain how Fig. X shows…"**: refer to the diagram. Name both the reactants and the products and say which is higher.
- **"Deduce the type of energy change"**: two marks, one for **exothermic/endothermic** and one for the **reason** (the levels, or the temperature change).
- **Temperature tables**: give the experiment number the question asks for. A "greatest change" can be a fall.
- **Completing a diagram**: "energy" on the axis, reactants on the left line, products on the right line. **Extended**: products line on the correct side, one upward E_a arrow from reactants to the top, one ΔH arrow from reactants to products. Each is usually a separate mark.
- **Bond energy calculations**: show each step on its own line with the numbers you added; later marks can follow through from an earlier slip. The last answer needs a **sign**, and units of kJ/mol (or kJ).
- **"Use the following steps"**: answer each step on its answer line: the value asked for there, not a running total with a letter in it.
- **Exothermic "in terms of bonds"**: say which is bigger, "energy released making bonds" or "energy needed breaking bonds". Both halves of the comparison are needed.
- **Reverse reaction**: same number, opposite sign, sign written.

Quick check

1. When 2 g of a solid is added to water the temperature changes from $21.0\ ^{\circ}\mathrm{C}$ to $27.5\ ^{\circ}\mathrm{C}$. Is the change exothermic or endothermic? Give a reason. **Extended:** what sign does ΔH have?
2. Three fuels are burned under $100\ \mathrm{cm}^3$ of water. Fuel X: 2.0 g burned, temperature rise $12.0\ ^{\circ}\mathrm{C}$. Fuel Y: 1.5 g, rise $10.5\ ^{\circ}\mathrm{C}$. Fuel Z: 3.0 g, rise $15.0\ ^{\circ}\mathrm{C}$. Which fuel releases the most energy per gram, and which gives the largest temperature rise?
3. **Extended:** on a reaction pathway diagram the reactants are at 50 kJ/mol, the top of the hump is at 210 kJ/mol and the products are at 120 kJ/mol. Find E_a and ΔH, and say whether the reaction is exothermic or endothermic.
4. **Extended:** calculate ΔH for $\mathrm{H_2 + Br_2 \rightarrow 2HBr}$, using H–H 436, Br–Br 193 and H–Br 366 kJ/mol. What is ΔH for $\mathrm{2HBr \rightarrow H_2 + Br_2}$?
5. Classify each as exothermic or endothermic: burning propane; heating calcium carbonate until it decomposes; neutralising sodium hydroxide with hydrochloric acid; photosynthesis.

Answers

1. The temperature **rises** by $27.5 - 21.0 = 6.5\ ^\circ\mathrm{C}$, so thermal energy was given out to the surroundings: **exothermic**. ΔH is **negative**.
2. Per gram: X $12.0 \div 2.0 = 6.0$, Y $10.5 \div 1.5 = 7.0$, Z $15.0 \div 3.0 = 5.0$ °C per gram. **Y** releases the most energy per gram. **Z** gives the largest temperature rise ($15.0\ ^\circ\mathrm{C}$), but only because more of it was burned.
3. $E_a = 210 - 50 = 160$ kJ/mol. $\Delta H = 120 - 50 = +70$ kJ/mol. The products are higher than the reactants: **endothermic**.
4. Broken: $436 + 193 = 629$ kJ. Made: $2 \times 366 = 732$ kJ. $\Delta H = 629 - 732 = -103$ kJ/mol (exothermic). The reverse reaction has $\Delta H = +103$ kJ/mol.
5. Burning propane: exothermic. Heating calcium carbonate: endothermic. Neutralisation: exothermic. Photosynthesis: endothermic.

Past questions to try

- **0620/33/M/J/25 Q6**: a table of temperatures, then labelling a reaction pathway diagram (Core).
- **0620/32/O/N/22 Q3**: the meaning of endothermic and comparing fuels per gram (Core).
- **0620/43/M/J/25 Q6**: part (e)(iii), a three-step bond energy calculation (Extended).
- **0620/41/O/N/25 Q4**: parts (a) and (b), drawing a diagram with E_a and ΔH, then finding an unknown bond energy (Extended).