

Cambridge AS & A Level · Chemistry 9701 · Notes

Chemical energetics

AS

Read online at pastopic.com/cambridge/chemistry-9701/notes/chemical-energetics-as

Last checked 30 September 2026

Read first: Atoms, molecules and stoichiometry, Chemical bonding

© 2026 Pastopic. Free for personal study. Please share the link, not the file.

What you need to know

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

1. **Say whether a reaction gives out or takes in energy.** Every reaction comes with an enthalpy change, ΔH . If heat is given out to the surroundings the reaction is **exothermic** and ΔH is negative; if heat is taken in it is **endothermic** and ΔH is positive.
2. **Draw and read a reaction pathway diagram**, showing the enthalpy change of the reaction and its activation energy.
3. **Define and use standard conditions** (298 K and 101 kPa, shown by the symbol $\ominus$) and the **standard enthalpy changes** of reaction (ΔH_r), formation (ΔH_f), combustion (ΔH_c) and neutralisation (ΔH_{neut}).
4. **Explain enthalpy changes with bonds**: breaking bonds takes energy in, making bonds gives energy out, and the balance of the two decides the sign of ΔH .
5. **Use bond energies** (always positive, because they are for breaking bonds) to calculate the enthalpy change of a reaction.
6. **Explain why some bond energies are exact and some are averages.**
7. **Work out enthalpy changes from experiments**, using $q = mc\Delta T$ and $\Delta H = -\frac{mc\Delta T}{n}$.
8. **Use Hess's law** to build simple energy cycles and do calculations with them, including finding enthalpy changes that cannot be measured directly and using bond energy data.

Lattice energy, Born–Haber cycles, hydration, entropy and Gibbs free energy are the A Level part of chemical energetics, not this topic.

Understand it

Enthalpy changes and their sign

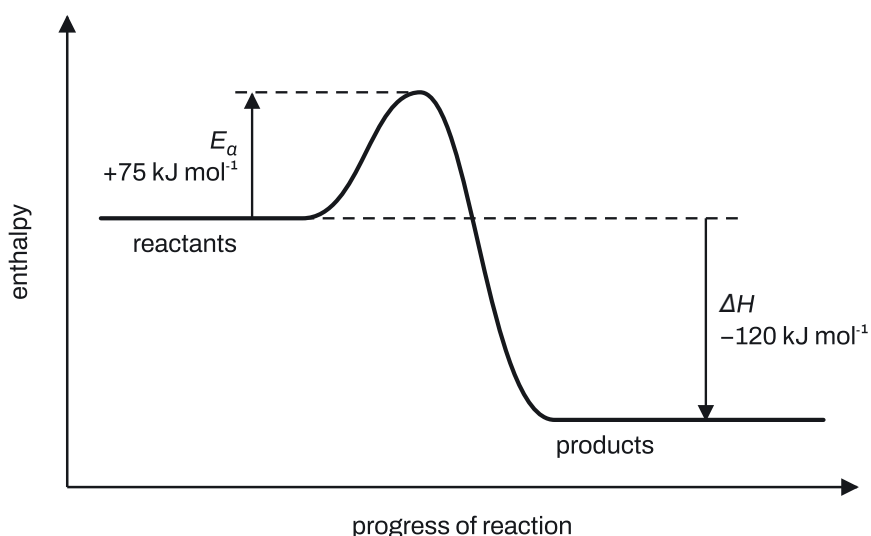
The **enthalpy**, H , of a set of chemicals is the energy stored in them. You can't measure it directly, but you can measure how much it **changes** in a reaction, because that energy goes into or comes out of the surroundings (the water, the beaker, the air) as heat. This is the **enthalpy change**, ΔH , measured in kJ mol^{-1} .

- **Exothermic:** the chemicals lose energy to the surroundings, so the surroundings get hotter and ΔH is **negative**. Combustion, neutralisation and the rusting of iron are exothermic.
- **Endothermic:** the chemicals take energy from the surroundings, so the surroundings get colder (or must be heated) and ΔH is **positive**. Thermal decomposition, such as breaking up a carbonate, is endothermic.

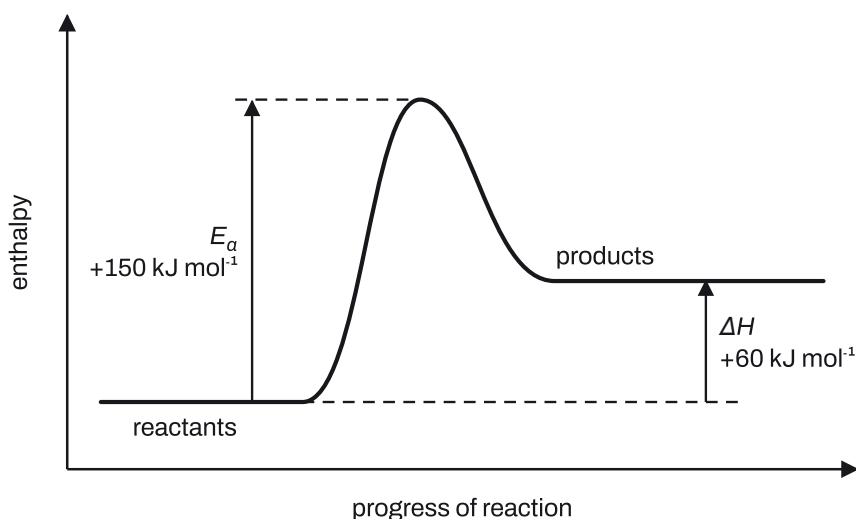
The sign is always from the chemicals' point of view: "the reaction mixture got hotter" means the chemicals gave energy away, so $\Delta H < 0$.

Reaction pathway diagrams

A reaction pathway diagram (also called an energy profile) plots enthalpy against the progress of the reaction. The reactants and products are flat lines at their enthalpy levels, and the reaction runs over a hump between them.



- ΔH is the vertical gap from the reactants to the products. Draw its arrow **from** the reactants **to** the products: down for exothermic, up for endothermic.
- **The activation energy, E_a ,** is the minimum energy colliding particles need to react. It is the gap from the reactants **up to the top of the hump**, so it is always positive.



The reverse reaction starts from the products and climbs the same hump, so

$$E_a(\text{reverse}) = E_a(\text{forward}) - \Delta H.$$

In the two diagrams: $75 - (-120) = 195 \text{ kJ mol}^{-1}$ for the reverse of the exothermic reaction, and $150 - 60 = 90 \text{ kJ mol}^{-1}$ for the reverse of the endothermic one.

If a reaction happens in **two steps**, the diagram has two humps with a dip between them (the intermediate). The step with the **higher** hump (larger activation energy, measured from the start of that step) is the **slower** step.

Standard enthalpy changes

The size of ΔH depends on temperature, pressure and the physical states of the chemicals, so values are compared under **standard conditions**: **298 K and 101 kPa**, with each substance in its **standard state** (its normal physical state under those conditions, such as $\text{H}_2\text{O}(\text{l})$, $\text{C}(\text{s})$ as graphite, $\text{Br}_2(\text{l})$, $\text{I}_2(\text{s})$). A standard value is written $\Delta H^\ominus$.

Name	Definition, in words	Example equation
Enthalpy change of reaction, $\Delta H_r^\ominus$	the enthalpy change when the amounts in the equation, as written, react under standard conditions	$2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$
Enthalpy change of formation, $\Delta H_f^\ominus$	the enthalpy change when one mole of a compound is formed from its elements in their standard states under standard conditions	$\text{Na}(\text{s}) + \frac{1}{2}\text{Cl}_2(\text{g}) \rightarrow \text{NaCl}(\text{s})$
Enthalpy change of combustion, $\Delta H_c^\ominus$	the enthalpy change when one mole of a substance is completely burned in oxygen under standard conditions	$\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$

Name	Definition, in words	Example equation
Enthalpy change of neutralisation, $\Delta H_{\text{neut}}^{\ominus}$	the enthalpy change when an acid and an alkali react to form one mole of water under standard conditions	$\text{H}^{+}(\text{aq}) + \text{OH}^{-}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$

Three things follow from these definitions:

- The ΔH_{f} of an **element in its standard state is zero** (forming $\text{O}_2(\text{g})$ from $\text{O}_2(\text{g})$ changes nothing).
- The equations for ΔH_{f} and ΔH_{c} must have exactly **1 mol** of the compound or fuel, so fractions such as $\frac{1}{2}\text{O}_2$ are normal.
- ΔH_{c} is always negative, and so is ΔH_{neut} . For any strong acid with any strong alkali, ΔH_{neut} is about -57 kJ mol^{-1} , because the reaction is always $\text{H}^{+} + \text{OH}^{-} \rightarrow \text{H}_2\text{O}$. With a **weak** acid (such as ethanoic acid) it is a little **less** negative: some of the energy released is used to finish ionising the acid as its H^{+} ions are removed.

Why energy changes: breaking and making bonds

In a reaction, bonds in the reactants break and new bonds form in the products.

- **Breaking a bond takes energy in** (endothermic): you have to pull the atoms apart against their attraction.
- **Making a bond gives energy out** (exothermic): the atoms fall together into a lower-energy arrangement.

A reaction is **exothermic** when more energy is given out making the new bonds than is taken in breaking the old ones, and **endothermic** when it is the other way round. So in an exothermic reaction between gases, the products' bonds are, in total, **stronger** than the reactants'.

The same idea covers changes of state: boiling and melting take energy in (intermolecular forces are overcome), and condensing and freezing give it out.

Bond energies

The **bond energy** of a bond is the energy needed to break one mole of that bond in gaseous molecules. It is always positive. For example, $E(\text{H-H}) = 436 \text{ kJ mol}^{-1}$ means $\text{H}_2(\text{g}) \rightarrow 2\text{H}(\text{g})$ has $\Delta H = +436 \text{ kJ mol}^{-1}$.

- **Exact** values are for bonds in one particular molecule, such as H-H (only in H_2) or H-Cl (only in HCl).
- **Average** values are for bonds that occur in many different compounds, such as C-H or C-C. A C-H bond in methane is not quite as strong as one in ethanol, because the neighbouring atoms change it slightly, so the table gives an average. The average C-H value is also found by splitting the energy to atomise $\text{CH}_4(\text{g})$ into $\text{C}(\text{g}) + 4\text{H}(\text{g})$ equally between the four bonds.

To find ΔH from bond energies:

$$\Delta H_{\text{r}} = \Sigma(\text{bonds broken}) - \Sigma(\text{bonds made}).$$

For $\text{CH}_2=\text{CH}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow \text{CH}_2\text{ClCH}_2\text{Cl}(\text{g})$, with $E(\text{C}=\text{C}) = 610$, $E(\text{Cl}-\text{Cl}) = 244$, $E(\text{C}-\text{C}) = 350$ and $E(\text{C}-\text{Cl}) = 340 \text{ kJ mol}^{-1}$: the four C-H bonds are unchanged, so leave them out. Broken: one C=C and one Cl-Cl, $610 + 244 = 854$. Made: one C-C and two C-Cl, $350 + 2(340) = 1030$. So $\Delta H = 854 - 1030 = -176 \text{ kJ mol}^{-1}$.

Answers from bond energies are only approximate, because average values are used, and bond energies only apply to gases (a liquid or solid also has intermolecular forces to overcome).

Measuring enthalpy changes

In a **calorimetry** experiment the reaction heats (or cools) a known mass of water or solution, and you measure the temperature change. The energy transferred is

$$q = mc\Delta T$$

where m is the mass of the **water or solution being heated** (in g), c is its specific heat capacity ($4.18 \text{ J g}^{-1} \text{ K}^{-1}$ for water; dilute solutions are treated as water, with a density of 1.00 g cm^{-3}), and ΔT is the temperature change. A change of 1°C is the same as a change of 1 K, so **don't add 273** to ΔT .

Then divide by the number of moles, n , that reacted, and give the sign:

$$\Delta H = -\frac{mc\Delta T}{n}$$

The minus sign is there because a temperature **rise** means an exothermic reaction. q is in J, so divide by 1000 to get kJ mol^{-1} .

- **Solution reactions** (neutralisation, displacement): m is the **total** volume of the mixed solutions, in g, not the mass of a solid added. n is the moles of the reactant that is **not** in excess (for ΔH_{neut} , the moles of water formed).
- **Combustion of a fuel**: a spirit burner heats water in a metal can. m is the mass of the **water**; n is the mass of fuel burned divided by its M_r .

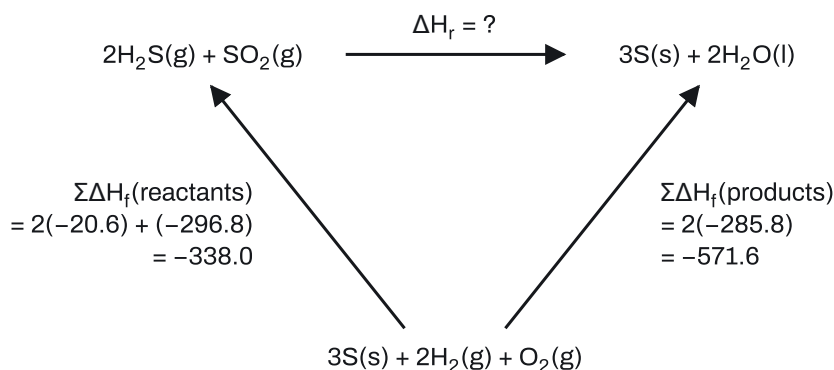
Experimental values are usually **less exothermic** than data-book values, because some heat escapes to the air and apparatus, and fuels may burn incompletely.

Hess's law

Hess's law: the enthalpy change of a reaction is the same whatever route is taken from the reactants to the products, as long as the starting and finishing conditions are the same. This is energy conservation: if two routes gave different values, you could go one way and back the other and create energy from nothing.

It lets you find enthalpy changes that are hard or impossible to measure, such as a reaction that is too slow, needs impossible conditions, or gives a mixture of products. Build a **cycle** with the reaction you want along the top and a second route through something you do have data for.

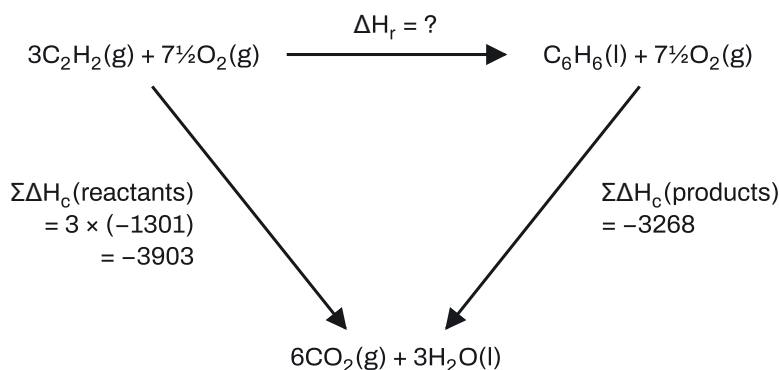
With enthalpies of formation, the second route goes through the **elements**. Both reactants and products can be made from the same elements, with arrows pointing **up** from the elements.



Going round the cycle, the route "elements $\rightarrow$ reactants $\rightarrow$ products" must equal "elements $\rightarrow$ products", so

$$\Delta H_r = \Sigma\Delta H_f(\text{products}) - \Sigma\Delta H_f(\text{reactants}).$$

With enthalpies of combustion, the second route goes through the **combustion products** (CO_2 and H_2O), with arrows pointing **down** to them.



This time the route "reactants $\rightarrow$ products $\rightarrow$ combustion products" must equal "reactants $\rightarrow$ combustion products", so

$$\Delta H_r = \Sigma\Delta H_c(\text{reactants}) - \Sigma\Delta H_c(\text{products}).$$

Notice the two formulas are the **opposite way round**. If you are unsure, draw the cycle and follow the arrows: going **with** an arrow, add its value; going **against** it, subtract.

With given equations, treat them like algebra: reverse an equation to change the sign of its ΔH , multiply an equation to multiply its ΔH , then add them so that the unwanted species cancel and you are left with the target equation.

Key facts and formulas

"Given" means it is in the 9701 Data Booklet. "Learn it" means you must remember it.

Formula	Status
Exothermic: $\Delta H < 0$; endothermic: $\Delta H > 0$	Learn it
Standard conditions: 298 K and 101 kPa, substances in standard states	Learn it
Definitions of ΔH_r , ΔH_f , ΔH_c , ΔH_{neut} (see the table above)	Learn it
ΔH_f of an element in its standard state = 0	Learn it
$E_a(\text{reverse}) = E_a(\text{forward}) - \Delta H$	Learn it
$\Delta H_r = \Sigma(\text{bonds broken}) - \Sigma(\text{bonds made})$	Learn it
Bond energy values, such as $E(\text{H-H}) = 436 \text{ kJ mol}^{-1}$	Given
$q = mc\Delta T$	Learn it
$c(\text{water}) = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$ ($4.18 \text{ kJ kg}^{-1} \text{ K}^{-1}$)	Given
$\Delta H = -\frac{mc\Delta T}{n}$	Learn it
Dilute solutions: density 1.00 g cm^{-3} , so 1 cm^3 has a mass of 1 g	Learn it
$\Delta H_r = \Sigma\Delta H_f(\text{products}) - \Sigma\Delta H_f(\text{reactants})$	Learn it
$\Delta H_r = \Sigma\Delta H_c(\text{reactants}) - \Sigma\Delta H_c(\text{products})$	Learn it
Strong acid + strong alkali: $\Delta H_{\text{neut}} \approx -57 \text{ kJ mol}^{-1}$	Learn it

How to do it

In the Paper 1 and Paper 2 questions we have tagged for this topic (104 questions, 2021–2025), the types came up this often. 12 of the 104 questions mix more than one type, so the counts add up to 119.

Question type	Questions
ΔH (or an unknown ΔH_f) from enthalpies of formation	23
Define a standard enthalpy change or choose its equation	23
ΔH from experimental results ($q = mc\Delta T$)	16
ΔH or an unknown bond energy from bond energies	13
Draw, label or read a reaction pathway diagram	13
Exothermic or endothermic, explained by bonds	10
ΔH from enthalpies of combustion	8
Combine given equations or read an energy cycle	8
Explain or predict neutralisation results	5

1. ΔH (or an unknown ΔH_f) from enthalpies of formation (23 questions)

1. **Write the balanced equation** with state symbols, if it isn't given.
2. **Multiply each ΔH_f by its coefficient** in the equation. Put in 0 for elements.
3. **Products minus reactants:** $\Delta H_r = \Sigma \Delta H_f(\text{products}) - \Sigma \Delta H_f(\text{reactants})$. Keep every value in brackets with its sign.
4. **Give the answer with its sign** and units, kJ mol^{-1} .

Variations you will meet:

- **An unknown ΔH_f .** Put the unknown in as x and solve. For the thermite reaction $\text{Fe}_2\text{O}_3(\text{s}) + 2\text{Al}(\text{s}) \rightarrow \text{Al}_2\text{O}_3(\text{s}) + 2\text{Fe}(\text{s})$, $\Delta H = -852 \text{ kJ mol}^{-1}$ and $\Delta H_f(\text{Al}_2\text{O}_3) = -1676 \text{ kJ mol}^{-1}$: $-852 = -1676 - x$, so $\Delta H_f(\text{Fe}_2\text{O}_3) = -824 \text{ kJ mol}^{-1}$.
- **ΔH_f from a reaction that makes 2 mol of the compound.** Divide by 2: ΔH_f is per **one** mole.
- **Multiple choice asking for the expression**, not the number: check the coefficients, which way round the subtraction goes, and that no element has a non-zero value.
- **Reaction schemes** where one step's ΔH is given and another is wanted: find the unknown step by going round the cycle.
- **Comparing ΔH_f values**, for example of two compounds linked by a reaction: write the reaction linking them and use the known ΔH .

2. Define a standard enthalpy change or choose its equation (23 questions)

1. **Say what is formed or used and how much:** "one mole of compound", "one mole of substance", "one mole of water".
2. **Say where it comes from:** "from its elements in their standard states" (ΔH_f), "burned completely in (excess) oxygen" (ΔH_c), "an acid reacting with an alkali" (ΔH_{neut}).
3. **Add "under standard conditions"** (298 K, 101 kPa) when the question says *standard*.

Choosing or writing the equation:

- Exactly **1 mol** of the named product (for ΔH_f) or of the fuel (for ΔH_c); use fractions elsewhere.
- Elements as they really are at 298 K: $\text{C}(\text{s})$, $\text{H}_2(\text{g})$, $\text{O}_2(\text{g})$, $\text{N}_2(\text{g})$, $\text{Cl}_2(\text{g})$, $\text{Br}_2(\text{l})$, $\text{I}_2(\text{s})$, metals (s), S(s). **Not** atoms such as $\text{C}(\text{g})$ or $\text{H}(\text{g})$, and **not** ions.
- The product in its standard state, for example ethanol and water as **liquids**.
- Standard conditions are **298 K**, not 298 °C, and **101 kPa**.
- ΔH for an equation that makes 2 mol (such as $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$) is an enthalpy change of **reaction**, equal to $2 \times \Delta H_f$.
- **Bond energy equations** have gases on both sides, for example $\text{HBr}(\text{g}) \rightarrow \text{H}(\text{g}) + \text{Br}(\text{g})$. For an average value such as C–H, the equation atomises the whole molecule, $\text{CH}_4(\text{g}) \rightarrow \text{C}(\text{g}) + 4\text{H}(\text{g})$, and the ΔH is divided by 4.
- A question may give the **enthalpy change of atomisation** of an element: the energy to make **1 mol of gaseous atoms** from the element in its standard state, such as $\frac{1}{2}\text{I}_2(\text{s}) \rightarrow \text{I}(\text{g})$. It is always positive.

3. ΔH from experimental results (16 questions)

1. **Energy transferred:** $q = mc\Delta T$, with m the mass of water or of the **whole** solution (total volume in $\text{cm}^3 =$ mass in g).
2. **Moles that reacted:** $n = c \times \frac{V}{1000}$ for a solution, or $n = \frac{\text{mass}}{M_r}$ for a fuel or solid. Use the reactant that is not in excess; for neutralisation, the moles of water formed.
3. $\Delta H = -\frac{q}{n}$, converted from J to kJ, with a **minus sign** if the temperature went up.
4. Give it to a sensible number of significant figures (usually 3) with units.

Variations you will meet:

- **A diprotic acid** such as H_2SO_4 : each mole gives 2 mol of H^+ , so it can make 2 mol of water.
- **Heat losses given as a percentage:** if only $x\%$ of the energy reached the water, the energy released is $q \times \frac{100}{x}$.
- **Predicting the temperature rise** from a known ΔH : work backwards, $\Delta T = \frac{n \times |\Delta H|}{mc}$. For 25.0 cm^3 of $0.800 \text{ mol dm}^{-3}$ HCl with 25.0 cm^3 of $0.800 \text{ mol dm}^{-3}$ KOH and $\Delta H_{\text{neut}} = -57.1 \text{ kJ mol}^{-1}$: $n = 0.0200 \text{ mol}$, $q = 0.0200 \times 57\,100 = 1142 \text{ J}$, $\Delta T = \frac{1142}{50.0 \times 4.18} = 5.5 \text{ K}$.
- **A heat capacity per cm^3 of solution** ($\text{J K}^{-1} \text{ cm}^{-3}$) instead of per gram: multiply by the volume instead of the mass.
- **Energy per gram** of fuel: divide the energy released by the mass burned, not by the moles.
- **Choosing the expression** (multiple choice): look for the total volume, ΔT without 273, and the moles of the limiting reactant.

4. ΔH or an unknown bond energy from bond energies (13 questions)

1. **Draw the displayed formulas** of everything in the equation, so you can count every bond.
2. **List the bonds broken** (reactants) and **the bonds made** (products), with how many of each. You may leave out bonds that appear unchanged on both sides, or count everything; both give the same answer.
3. $\Delta H = \Sigma(\text{broken}) - \Sigma(\text{made})$.

Variations you will meet:

- **Find an unknown bond energy:** put it in as x and solve, remembering how many of that bond there are.
- **Mixed with ΔH_f or atomisation data**, for example to find a mean bond energy: build a cycle through the **gaseous atoms** (Example 4 below).
- **Why the answer differs from the data-book value:** bond energies are averages, and they apply only to gases.
- **Statements about bond energies:** exact values for a bond found in one molecule, averages for bonds found in many compounds.

5. Draw, label or read a reaction pathway diagram (13 questions)

1. **Label the axes:** enthalpy (or energy) up the side, progress of reaction along the bottom.
2. **Draw reactants and products** as flat lines: products **lower** for exothermic, **higher** for endothermic. Write their names or formulas on the lines.
3. **Draw the hump** between them, higher than both.
4. **Arrows:** ΔH from the reactants' level to the products' level (arrowhead at the products); E_a from the reactants' level up to the top of the hump.

Variations you will meet:

- **Choosing the right diagram** from given ΔH and E_a : check the sign of ΔH and the ratio of E_a to ΔH (for example " E_a is twice the size of ΔH ").
- **Reverse activation energy:** $E_a(\text{reverse}) = E_a(\text{forward}) - \Delta H$.
- **Two-step reactions:** two humps; the smaller hump is the faster step.
- **A diagram drawn for more than one mole:** for example, a diagram for $2\text{NaOH} + \text{H}_2\text{SO}_4$ shows the energy for 2 mol of water, so ΔH_{neut} is half of it.

6. Exothermic or endothermic, explained by bonds (10 questions)

1. Breaking bonds is **endothermic**; making bonds is **exothermic**.
2. **Exothermic** means more energy is **given out forming** bonds than is **taken in breaking** them (and the other way round for endothermic).
3. For gases, a negative ΔH means the products' total bond energy is **greater** than the reactants'.

Signs to know: combustion, neutralisation and oxidation of metals (such as iron in a hand warmer) are exothermic; atomisation, bond breaking and thermal decomposition are endothermic; condensing and freezing are exothermic; melting and boiling are endothermic. Enthalpy changes of reaction in general can have either sign; only some kinds, such as combustion and neutralisation, are always negative.

7. ΔH from enthalpies of combustion (8 questions)

1. **Multiply each ΔH_c by its coefficient.**
2. **Reactants minus products:** $\Delta H_r = \Sigma \Delta H_c(\text{reactants}) - \Sigma \Delta H_c(\text{products})$.
3. **To find ΔH_f of a compound**, the "reactants" are its elements: for propane, $3\text{C}(\text{s}) + 4\text{H}_2(\text{g}) \rightarrow \text{C}_3\text{H}_8(\text{g})$, so with ΔH_c values of -394 (C), -286 (H_2) and -2220 (C_3H_8) kJ mol^{-1} : $\Delta H_f = 3(-394) + 4(-286) - (-2220) = -106 \text{ kJ mol}^{-1}$.

Variations you will meet:

- **A two-step scheme** (reaction 1 then reaction 2, with the overall change given): find one step from combustion data, then the other by going round the scheme.
- **Working backwards** to a ΔH_c from ΔH_f values: write the combustion equation and use products minus reactants (type 1).

- **Multiple choice asking for the expression:** check the direction of the subtraction.

8. Combine given equations or read an energy cycle (8 questions)

1. **Write the target equation** and mark which side each species must be on.
2. **Take each given equation** in turn: reverse it if its species are on the wrong side (change the sign of ΔH), and multiply it so the amounts match (multiply ΔH too).
3. **Add** the adjusted equations and their ΔH values. Check everything unwanted cancels.

For example, to find $\Delta H_f(\text{CO})$ from $\text{C(s)} + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$ (-394 kJ mol^{-1}) and $\text{CO(g)} + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$ (-283 kJ mol^{-1}): keep the first, reverse the second, and add: $-394 + 283 = -111 \text{ kJ mol}^{-1}$.

Variations you will meet:

- **Reading a drawn cycle** (multiple choice): identify what each arrow stands for, and whether it is added or subtracted going round.
- **Choosing which data would give ΔH :** formation data, combustion data or bond energies each work only if every species in the equation has a value (bond energies only for gases).
- **Dividing at the end:** if the target is "per mole" of something that appears 2 or 4 times, divide the total.

9. Explain or predict neutralisation results (5 questions)

- **Weak acid** (ethanoic acid) with a strong alkali: ΔH_{neut} is **less negative** than about -57 kJ mol^{-1} , because the weak acid is only partly ionised and some energy is used to ionise the rest as the reaction goes on.
- **Changing amounts:** if both volumes are doubled at the same concentrations, twice the energy heats twice the mass, so ΔT is **the same**. If both concentrations are doubled at the same volumes, ΔT **doubles**.
- **Diprotic acid:** H_2SO_4 gives twice as much H^+ per mole, so with an excess of it the alkali is the limiting reactant.
- **A reaction unsuitable for measuring ΔH_{neut} :** if a precipitate also forms (for example barium hydroxide with sulfuric acid gives solid BaSO_4), its enthalpy change is added in and the result is not just neutralisation.

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

Some standard enthalpy changes of formation are given.

Substance	$\Delta H_f^\ominus / \text{kJ mol}^{-1}$
$\text{H}_2\text{S}(\text{g})$	-20.6
$\text{SO}_2(\text{g})$	-296.8
$\text{H}_2\text{O}(\text{l})$	-285.8

(a) Write the equation, with state symbols, for the reaction whose enthalpy change is the standard enthalpy change of formation of hydrogen sulfide. **[1]**

(b) Calculate the standard enthalpy change for the reaction $2\text{H}_2\text{S}(\text{g}) + \text{SO}_2(\text{g}) \rightarrow 3\text{S}(\text{s}) + 2\text{H}_2\text{O}(\text{l})$. **[2]**

(c) Explain why no value for sulfur is needed in (b). **[1]**

(a) $\text{H}_2(\text{g}) + \text{S}(\text{s}) \rightarrow \text{H}_2\text{S}(\text{g})$ **B1** (1 mol of H_2S , elements in their standard states, all state symbols correct).

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

$$\Delta H = 2(-285.8) - [2(-20.6) + (-296.8)] = -571.6 - (-338.0)$$

M1 for the correctly weighted expression.

$$\Delta H = -233.6 \text{ kJ mol}^{-1}$$

A1 (the sign is needed). The first diagram of "Hess's law" in "Understand it" is this cycle.

(c) Sulfur is an element in its standard state, so its ΔH_f is zero. **B1**

Example 2

Benzene can, in principle, be made from ethyne: $3\text{C}_2\text{H}_2(\text{g}) \rightarrow \text{C}_6\text{H}_6(\text{l})$.

Substance	$\Delta H_c^\ominus / \text{kJ mol}^{-1}$
$\text{C}_2\text{H}_2(\text{g})$	-1301
$\text{C}_6\text{H}_6(\text{l})$	-3268

(a) Define standard enthalpy change of combustion. **[2]**

(b) Calculate the standard enthalpy change for the conversion of ethyne into benzene. **[2]**

(c) Suggest why this enthalpy change cannot easily be measured directly. **[1]**

(a) The enthalpy change when one mole of a substance is burned completely in oxygen **B1**, under standard conditions (298 K, 101 kPa) **B1**.

(b) Both sides burn to $6\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l})$, so

$$\Delta H = 3(-1301) - (-3268)$$

M1 for reactants minus products, with the 3.

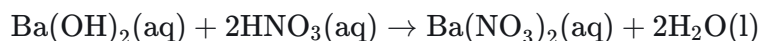
$$\Delta H = -3903 + 3268 = -635 \text{ kJ mol}^{-1}$$

A1. The second Hess's law diagram in "Understand it" shows this cycle.

(c) The reaction does not happen cleanly under standard conditions: it is very slow, or it gives other products as well, so the heat measured would not be for this reaction alone. **B1**

Example 3

A student measures the enthalpy change of neutralisation for nitric acid with barium hydroxide.



40.0 cm³ of 0.500 mol dm⁻³ Ba(OH)₂(aq) is mixed with 60.0 cm³ of 0.800 mol dm⁻³ HNO₃(aq) in a polystyrene cup. The temperature rises from 19.6 °C to 25.0 °C.

(a) Define enthalpy change of neutralisation. **[1]**

(b) Calculate the enthalpy change of neutralisation. Assume the solution has a density of 1.00 g cm⁻³ and a specific heat capacity of 4.18 J g⁻¹ K⁻¹. **[4]**

(c) The experiment is repeated with 60.0 cm³ of 0.800 mol dm⁻³ ethanoic acid instead of nitric acid. Predict how the temperature rise compares, and explain your answer. **[2]**

(a) The enthalpy change when an acid and an alkali react to form one mole of water. **B1**

(b) $q = mc\Delta T = 100.0 \times 4.18 \times 5.4 = 2257.2 \text{ J}$ **M1** (mass = total volume 40.0 + 60.0; $\Delta T = 5.4 \text{ K}$).

$$n(\text{OH}^-) = 2 \times 0.500 \times \frac{40.0}{1000} = 0.0400 \text{ mol}; n(\text{H}^+) = 0.800 \times \frac{60.0}{1000} = 0.0480 \text{ mol}.$$

So OH⁻ is the limiting reactant and 0.0400 mol of water forms. **M1**

$$\Delta H = -\frac{2257.2}{0.0400} = -56\,430 \text{ J mol}^{-1}$$

M1 for dividing by the moles of water.

$$\Delta H_{\text{neut}} = -56.4 \text{ kJ mol}^{-1}$$

A1 (3 s.f., negative sign, kJ mol⁻¹).

(c) The temperature rise is smaller. **B1** Ethanoic acid is a weak acid and is only partly ionised, so some of the energy released is used to ionise it fully. **B1**

Example 4

Some data are given.

Data	Value / kJ mol ⁻¹
$\Delta H_f^\ominus$ of CF ₄ (g)	−933
Enthalpy change of atomisation of carbon, C(s) → C(g)	+717
Bond energy F–F	158
Bond energy C–H	410
Bond energy H–F	562

(a) Calculate the mean C–F bond energy in CF₄. **[3]**

(b) Use your answer to (a) and the other bond energies to calculate ΔH for CH₄(g) + 4F₂(g) → CF₄(g) + 4HF(g). **[2]**

(c) Using standard enthalpy changes of formation, the value for this reaction is −1950 kJ mol⁻¹. Suggest why it differs from your answer to (b). **[1]**

(a) Build a cycle from the elements, through gaseous atoms, to CF₄(g):

- C(s) + 2F₂(g) → C(g) + 4F(g): 717 + 2(158) = +1033 kJ mol⁻¹ **M1**
- C(g) + 4F(g) → CF₄(g): −4E(C–F), since four bonds form.

By Hess's law, $\Delta H_f = 1033 - 4E(\text{C–F})$, so $-933 = 1033 - 4E$. **M1**

$$E(\text{C–F}) = \frac{1033 + 933}{4} = \frac{1966}{4} = 491.5 \text{ kJ mol}^{-1}$$

A1

(b) Broken: 4 C–H + 4 F–F = 4(410) + 4(158) = 2272. Made: 4 C–F + 4 H–F = 4(491.5) + 4(562) = 4214. **M1**

$$\Delta H = 2272 - 4214 = -1942 \text{ kJ mol}^{-1}$$

A1

(c) The C–H bond energy is an average value from many compounds, not the exact value for methane. **B1**

Common mistakes

- **Adding 273 to the temperature change.** A rise of 5.4 °C is a rise of 5.4 K. Adding 273 to ΔT is a favourite wrong option in multiple choice.
- **Using the wrong mass in $q = mc\Delta T$.** It is the mass of the water or solution being heated: the **total** volume of both solutions, not one of them, and not the mass of the solid or the fuel.
- **Forgetting to divide by moles, or dividing by the wrong moles.** Dividing q by 1000 turns J into kJ, but you still need "per mole". Use the limiting reactant, and for neutralisation the moles of **water** (twice the acid for H₂SO₄).

- **Losing the sign.** A temperature rise means a **negative** ΔH . Always write the sign, including "+" for endothermic.
- **Definitions without "one mole".** "The energy when a compound is formed from its elements" misses the point that it is for 1 mol, and that the elements are in their standard states. Some candidates even write compounds forming from ions or atoms.
- **Wrong state symbols in formation equations.** Carbon is C(s), iodine I₂(s), water H₂O(l); state symbols are often missing or wrong.
- **Subtracting the wrong way round.** Formation data: products minus reactants. Combustion data: reactants minus products. Drawing the cycle prevents this; examiners find that candidates who sketch one do better.
- **Missing the coefficients.** Every ΔH_f or bond energy must be multiplied by how many there are (for example 2H₂O counts twice). Leaving one out is the commonest arithmetic error.
- **Bonds broken minus bonds made, the wrong way round**, or using O–O instead of O=O, or counting a C–C bond twice.
- **Mislabelling pathway diagrams.** E_a is measured from the **reactants** to the top of the hump, not from the products; the ΔH arrow goes from reactants to products.
- **Forgetting to halve.** An enthalpy change for "per mole of water" or "per mole of compound" from an equation that makes 2 mol needs dividing by 2.

How to get the marks

- **Calculations (usually 2–3 marks on Paper 2):** the first mark is for the correctly set-up expression, such as $2(-285.8) - [2(-20.6) + (-296.8)]$, and the last for the final value **with its sign and units**. Write the expression down before working it out; if the arithmetic goes wrong you keep the method mark.
- **Experimental calculations** earn marks separately for q , for the moles, and for ΔH with sign and units. Wrong earlier answers are carried forward, so finish every part.
- **Definitions** are marked point by point: one mole of what, formed from or reacting with what, standard conditions. Learn them word-perfect in meaning, not copied phrasing.
- **"Write an equation"** for a standard enthalpy change means balanced, 1 mol of the right species and correct state symbols.
- **"Draw a reaction pathway diagram":** labelled axes, reactants and products at the right levels, a hump, and labelled arrows for ΔH and E_a pointing the right way.
- **"Explain" with bonds:** compare the energy **taken in** breaking bonds with the energy **given out** making them; say which is larger.
- **Significant figures:** give the final answer to 3 s.f. unless the data suggest otherwise; don't round in the middle of a calculation.
- **Multiple choice:** work the answer out yourself before looking at the options; the wrong options are the usual slips (no coefficient, wrong sign, +273, wrong mass).

Quick check

1. Write equations, with state symbols, for (a) the standard enthalpy change of formation of potassium nitrate, $\text{KNO}_3(\text{s})$, and (b) the standard enthalpy change of combustion of propan-1-ol, $\text{C}_3\text{H}_7\text{OH}(\text{l})$.
2. A reaction has $\Delta H = -58 \text{ kJ mol}^{-1}$ and an activation energy of $+134 \text{ kJ mol}^{-1}$. Sketch its reaction pathway diagram and find the activation energy of the reverse reaction.
3. Burning 0.480 g of methanol, CH_3OH ($M_r = 32.0$), heats 200 g of water by 11.5 K. Calculate the enthalpy change of combustion of methanol from these results, and give one reason why it is less exothermic than the data-book value.
4. $\text{S}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g})$, $\Delta H = -297 \text{ kJ mol}^{-1}$, and $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{SO}_3(\text{g})$, $\Delta H = -198 \text{ kJ mol}^{-1}$. Find the enthalpy change of formation of $\text{SO}_3(\text{g})$.
5. (a) Use the bond energies H–H 436, Br–Br 193 and H–Br 366 kJ mol^{-1} to calculate ΔH for $\text{H}_2(\text{g}) + \text{Br}_2(\text{g}) \rightarrow 2\text{HBr}(\text{g})$. (b) The H–Br bond energy is an exact value but the C–Br bond energy is an average. Explain why.

Answers

1. (a) $\text{K(s)} + \frac{1}{2}\text{N}_2(\text{g}) + \frac{3}{2}\text{O}_2(\text{g}) \rightarrow \text{KNO}_3(\text{s})$. (b) $\text{C}_3\text{H}_7\text{OH(l)} + \frac{9}{2}\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O(l)}$.
2. Products drawn 58 kJ mol^{-1} **below** the reactants, a hump 134 kJ mol^{-1} above the reactants, ΔH arrow pointing down, E_a arrow from the reactants up to the top. $E_a(\text{reverse}) = 134 - (-58) = 192 \text{ kJ mol}^{-1}$.
3. $q = 200 \times 4.18 \times 11.5 = 9614 \text{ J}$; $n = \frac{0.480}{32.0} = 0.0150 \text{ mol}$; $\Delta H_c = -\frac{9.614}{0.0150} = -641 \text{ kJ mol}^{-1}$ (3 s.f.).
It is less exothermic because heat is lost to the air and the apparatus (or the methanol burns incompletely).
4. Target: $\text{S(s)} + \frac{3}{2}\text{O}_2(\text{g}) \rightarrow \text{SO}_3(\text{g})$. Use the first equation as it is and half of the second: $\Delta H_f = -297 + \frac{1}{2}(-198) = -396 \text{ kJ mol}^{-1}$.
5. (a) Broken: $436 + 193 = 629$; made: $2 \times 366 = 732$; $\Delta H = 629 - 732 = -103 \text{ kJ mol}^{-1}$. (b) The H-Br bond is only found in the HBr molecule, so its value is measured for that one molecule. C-Br bonds occur in many different compounds, where the other atoms change their strength slightly, so the value is averaged over them.

Past questions to try

- **9701/21/M/J/24 Q4**: the definition of ΔH_f , its equation and a formation-data calculation.
- **9701/21/M/J/23 Q2**: a neutralisation experiment with sulfuric acid, and why Ba(OH)_2 with H_2SO_4 can't be used.
- **9701/23/O/N/25 Q3**: a two-step scheme using enthalpies of combustion (part (d)).
- **9701/22/M/J/25 Q3**: enthalpy change of a hydrogenation from bond energies (part (c)).