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IB DIPLOMA CHEMISTRY · FIRST ASSESSMENT 2025

Reactivity 1.1
Measuring Enthalpy Changes

Booklet 1 · Understandings 1.1.1 – 1.1.4

Heat and temperature, system and surroundings, exothermic and endothermic reactions, energy profiles, standard enthalpy changes and calorimetry calculations with $Q = mc\Delta T$

Complete revision notes, worked examples, model answers & practice questions · Standard Level & Higher Level

How the level tags work in this booklet

SL + HL Every understanding in Reactivity 1.1 is common to both levels, so the whole booklet is core material that every student must know.

★ marks harder, multi-step practice questions of the kind that separate grade 6 from grade 7. HL students meet these ideas again in Reactivity 1.2 (Hess's law) and Reactivity 1.4 (entropy and Gibbs energy).

Inside this booklet

1. What the syllabus actually asks of you
2. 1.1.1 — Heat, temperature and energy conservation
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1 What the syllabus actually asks of you

Reactivity 1.1 opens the "What drives chemical reactions?" strand. Its job is to give you the language of energy change (system, surroundings, enthalpy, exothermic, endothermic) and one key experimental skill: turning a temperature change into an enthalpy change. Almost every energetics question in Papers 1 and 2 starts here, and the calculation in understanding 1.1.4 is one of the most frequently examined skills in the whole course.

Understanding	What it says	What you must be able to do	Level
1.1.1 Energy transfer	Chemical reactions involve a transfer of energy between the system and the surroundings, while total energy is conserved.	Understand the difference between **heat** and **temperature**. Identify the system and the surroundings, and state that energy lost by one is gained by the other.	SL+HL
1.1.2 Exo and endo	Reactions are described as endothermic or exothermic, depending on the direction of energy transfer between the system and the surroundings.	Understand the temperature change (decrease or increase) that accompanies endothermic and exothermic reactions, respectively.	SL+HL
1.1.3 Stability	The relative stability of reactants and products determines whether reactions are endothermic or exothermic.	Sketch and interpret potential energy profiles for endothermic and exothermic reactions. Axes are labelled **reaction coordinate** (x) and **potential energy** (y).	SL+HL
1.1.4 Measuring ΔH	The standard enthalpy change for a chemical reaction, $\Delta H^{\ominus}$, refers to the heat transferred at constant pressure under standard conditions and states. It can be determined from the change in temperature of a pure substance.	Apply $\mathbf{Q = mc\Delta T}$ and $\mathbf{\Delta H = -Q/n}$ to calculate the enthalpy change of a reaction. The units of $\Delta H^{\ominus}$ are kJ mol^{-1}. Evaluate calorimetry experiments and explain why the measured temperature change is usually smaller than expected.	SL+HL

Linking questions and connections for 1.1.1 – 1.1.4

- **Structure 1.1:** temperature is a measure of the average kinetic energy of particles. → section 2
- **Structure 1.4 and Reactivity 2.1:** amounts in mol, n = m/M, n = cV, and identifying the limiting reactant. → sections 7 and 8
- **Reactivity 2.2:** activation energy, E_a, and why some energetically favourable reactions are slow. → section 4
- **Reactivity 1.2:** bond enthalpies and Hess's law explain *why* a reaction is exothermic; the H_2O_2 diagram in section 4 is your first Hess cycle.
- **Reactivity 1.4:** entropy explains why some endothermic reactions are spontaneous. → section 4.3
- **Tool 1, Inquiry 1–3:** how can enthalpy changes be investigated experimentally, and why do calorimetry experiments measure a smaller temperature change than expected? → sections 7 and 8

What the data booklet gives you — and what it does not

The data booklet gives the equation $\mathbf{Q = mc\Delta T}$, the specific heat capacity of water, $\mathbf{c = 4.18\ J\ g^{-1}\ K^{-1}}$, and a table of standard enthalpies of combustion that you compare your experimental values with. Learn $\mathbf{\Delta H = -Q/n}$, the definitions, the standard conditions and the sign convention yourself: they are tested directly and you should never depend on looking them up.

2 Reactivity 1.1.1 — Heat, temperature and energy conservation SL + HL

2.1 Energy, work and heat

Energy is the capacity to do **work**, that is, to move an object against an opposing force. Its SI unit is the joule (J): you do 1 J of work when you push with a force of 1 N through a distance of 1 m. Energy can be stored, for example as the **chemical potential energy** held in bonds and intermolecular forces, and it can be transferred in several ways: as heat, as light, as sound or electrically. This topic is about the **heat** transfers that accompany chemical reactions; electrical energy transfers appear in Reactivity 3.2.

Definitions to learn

Heat is energy transferred from one place to another as a result of a temperature difference. It always flows from the hotter object to the colder one.

Temperature is a measure of the **average kinetic energy** of the particles in a substance. The absolute temperature in kelvin is directly proportional to the average kinetic energy of the particles.

The distinction matters because the two behave differently. Temperature is a **property** of an object; heat is a **process**, energy on the move. When heat is added to a substance, its particles move faster on average, so its temperature rises, and the extra kinetic energy is spread among all the particles present.

	Heat	Temperature
What it is	a process: energy transferred because of a temperature difference	a property: a measure of the average kinetic energy of the particles
Depends on amount?	yes: more particles need more energy for the same temperature rise	no: a cup and a bath of water can both be at 20 °C
Unit	joule, J (or kJ)	kelvin, K (or °C)
How it is found	calculated from mass, specific heat capacity and temperature change	measured directly with a thermometer or probe

A useful illustration: far more heat is needed to raise the temperature of a swimming pool by 1 K than to raise the water in a kettle by 1 K, even though both start at the same temperature. The pool contains many more water molecules, and each one must gain the same extra kinetic energy, so the total energy transferred is much larger.

2.2 Kelvin and Celsius

The two scales have intervals of exactly the same size: $T / K = T / °C + 273.15$. A temperature **change** is therefore the same number on both scales. A rise from 300 K to 310 K is a rise of 10 K, which is also a rise of 10 °C. You never add 273 to a temperature change, and ΔT may be quoted in K or °C in any calculation.

2.3 System and surroundings

Chemical changes happen in test tubes, polystyrene cups, industrial reactors and living cells. In each case it helps to separate the **system**, the part we are interested in (usually the reacting chemicals), from the **surroundings**, which in principle is everything else in the universe. In practice the surroundings are the solvent, the container, the thermometer and the air around them.

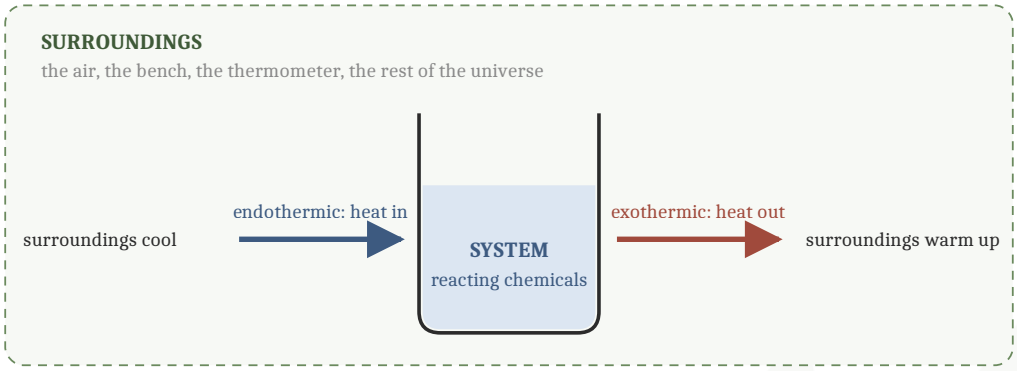


Figure 1 — The system is the reaction of interest; the surroundings are everything else. Heat flowing out of the system warms the surroundings (exothermic); heat flowing in cools them (endothermic).

Type of system	Exchanges energy?	Exchanges matter?	Example
Open	yes	yes	magnesium reacting with acid in an open beaker: heat and hydrogen gas both escape
Closed	yes	no	a sealed flask: it warms the air outside but no gas escapes
Isolated	no	no	an ideal sealed, perfectly insulated flask (approximated by a lidded vacuum flask)

Most reactions you carry out take place in **open** systems. Whatever the type, the law of conservation of energy applies: energy cannot be created or destroyed, so any energy lost by the system is gained by the surroundings, and vice versa.

Worked example — conservation of energy in a simple transfer

Q. A hot block of iron at temperature T_H is placed in contact with a cooler block of iron at T_C. The two blocks are insulated from everything else. Describe what happens.

A. Heat flows from the hotter block to the cooler block because of the temperature difference. The particles of the hot block lose kinetic energy on average, so its temperature falls; the particles of the cool block gain kinetic energy, so its temperature rises. Transfer continues until both blocks reach the **same temperature**, when there is no longer a temperature difference to drive it. Because the pair is insulated, the **total energy** of the two blocks stays constant: the energy lost by one equals the energy gained by the other.

Model answer: "Distinguish between heat and temperature." (2 marks)

Heat is a form of energy transfer that occurs because of a temperature difference, measured in joules **[1]**. Temperature is a measure of the average kinetic energy of the particles, measured in kelvin (or °C) **[1]**.

Examiner traps in 1.1.1

- "Heat is how hot something is" describes temperature, not heat, and scores nothing.
- Temperature measures the **average** kinetic energy, not the total. A large object at the same temperature has more total energy but the same average.
- Cold does not flow. A cold pack feels cold because heat flows **from your hand** into it.
- Never convert ΔT by adding 273: a change of 15 °C is a change of 15 K.

3 Reactivity 1.1.2 — Exothermic and endothermic reactions SL + HL

3.1 Enthalpy and the sign of ΔH

The chemical potential energy stored in a system is called its **enthalpy**, H, from the Greek for "heat inside". It is held in the chemical bonds and intermolecular forces of the substances present. We cannot measure the absolute enthalpy of a substance, but we can measure the **change** in enthalpy, ΔH, from the heat transferred when the reaction happens at constant pressure.

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

ΔH is negative when the system loses enthalpy, positive when it gains enthalpy

Exothermic reaction

Heat is transferred **from the system to the surroundings**. The products have less enthalpy than the reactants, so **ΔH is negative**. The surroundings, including any solvent, **warm up**.

Endothermic reaction

Heat is transferred **from the surroundings to the system**. The products have more enthalpy than the reactants, so **ΔH is positive**. The surroundings **cool down**.

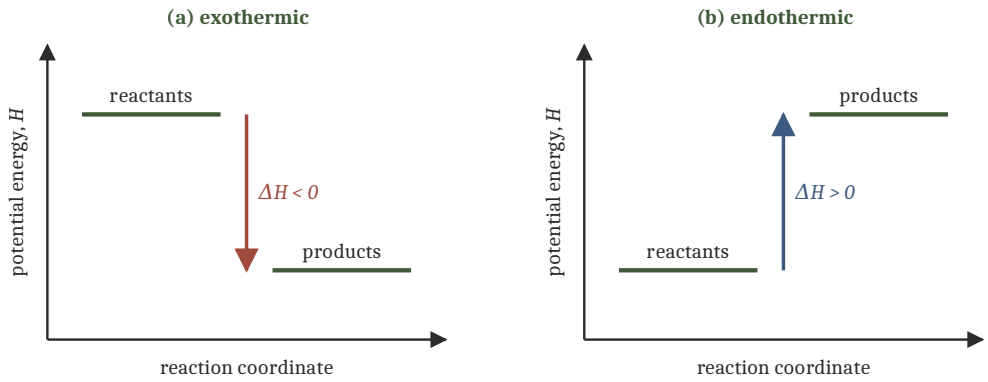


Figure 2 — Enthalpy level diagrams. The arrow always runs from reactants to products: downwards (ΔH < 0) for an exothermic reaction, upwards (ΔH > 0) for an endothermic reaction.

3.2 Temperature changes during reactions in solution

When a reaction happens in aqueous solution, the water acts as the surroundings that the thermometer actually measures. In an **exothermic** reaction the heat released by the reacting chemicals is first passed to the water, whose molecules gain kinetic energy, so the **temperature rises**. Over time the warm water passes this heat to the wider surroundings (the air, the bench) and slowly returns to room temperature. The net result is a transfer of heat from the system to the surroundings and a decrease in the enthalpy of the system.

In an **endothermic** reaction the reaction takes the heat it needs from the water, so the **temperature falls**. Later, heat flows back in from the room and the temperature returns to its starting value. A dramatic example is the reaction between solid ammonium thiocyanate, NH_4SCN, and solid barium hydroxide octahydrate, $Ba(OH)_2{\cdot}8H_2O$: it absorbs so much heat that a wet beaker can freeze to the bench.

> **The idea in one line**
>
> The thermometer reads the **surroundings**. Temperature up means the surroundings gained heat, so the system lost it: **exothermic, ΔH negative**. Temperature down means **endothermic, ΔH positive**.

3.3 Common examples

Exothermic (ΔH negative)	Endothermic (ΔH positive)
combustion of fuels: CH_4, alcohols, H_2, coal	photosynthesis: $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$
neutralization of acids by bases	thermal decomposition, e.g. $CaCO_3 \rightarrow CaO + CO_2$
reactive metals with acids or water; displacement, e.g. $Zn + Cu^{2+}$	$NH_4SCN(s) + Ba(OH)_2{\cdot}8H_2O(s)$
respiration	dissolving many ammonium salts, e.g. NH_4NO_3, NH_4Cl, and KNO_3
physical changes: condensing, freezing	physical changes: melting, boiling, evaporating, subliming
forming bonds (in any process)	breaking bonds (in any process); $N_2 + O_2 \rightarrow 2NO$

Bond breaking always absorbs energy and bond making always releases it. Whether a reaction is exothermic overall depends on the balance between the two, which you will calculate in Reactivity 1.2.

Worked example — classifying from observations

Observation	Reasoning	Verdict
Magnesium ribbon added to dilute HCl: the test tube becomes warm.	Surroundings gain heat, so the system loses heat.	exothermic, $\Delta H < 0$
Ammonium chloride stirred into water: the beaker feels cold.	Heat flows from the water (and your hand) into the system.	endothermic, $\Delta H > 0$
Ethanol rubbed on skin evaporates: the skin feels cool.	Evaporation absorbs heat; the skin is the surroundings and loses heat.	endothermic, $\Delta H > 0$
Steam condenses on a cold window, which warms slightly.	Condensation forms intermolecular forces and releases heat.	exothermic, $\Delta H < 0$

Model answer: "When ammonium nitrate dissolves in water the temperature falls from 21.0 °C to 15.5 °C. Deduce the sign of ΔH and explain your answer." (2 marks)

ΔH is positive; the process is endothermic **[1]**. The dissolving absorbs heat from the water (the surroundings), so the average kinetic energy of the water molecules, and therefore the temperature, decreases **[1]**.

Examiner traps in 1.1.2

- "The temperature went down, so heat was lost and ΔH is negative" is the classic wrong answer. A temperature fall means the system **gained** heat from the water.
- Always state the **direction** of transfer: "heat is released **from the system to the surroundings**", not just "heat is released".
- Exothermic does not mean fast, and it does not mean the reaction needs no energy to start (see activation energy, section 4).
- Do not call melting or evaporating "reactions"; they are physical changes, but they still have ΔH values and signs.

4 Reactivity 1.1.3 — Energetic stability and energy profiles SL + HL

4.1 Stability and the direction of change

Change has a natural direction. Someone who slips on a ladder falls down, not up, because the lower position has less potential energy and is more stable. Chemical systems behave in a similar way: most reactions that happen readily are ones that **lower** the potential energy (enthalpy) of the system. A mixture of carbon dioxide and water is lower in enthalpy, and so more stable, than the methane and oxygen it came from.

Definition to learn

The **lower the enthalpy** of a substance or mixture, the **more energetically stable** it is. In an exothermic reaction the products are more stable than the reactants; in an endothermic reaction the reactants are more stable than the products.

Stability is always a **relative** term: it only makes sense to say A is more stable *than* B. Hydrogen peroxide is more stable than the hydrogen and oxygen it can be made from, but less stable than the water and oxygen it decomposes into.

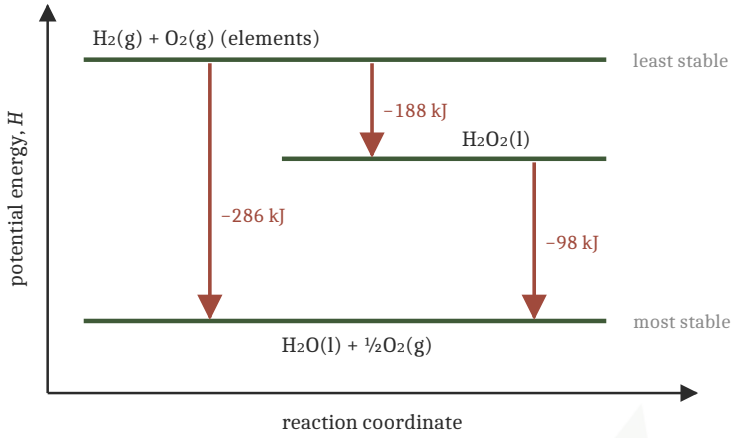


Figure 3 — Hydrogen peroxide is stable relative to its elements but unstable relative to water and oxygen. Notice that −286 = −188 + (−98): the enthalpy change does not depend on the route, which is Hess's law (Reactivity 1.2). Values in kJ per mole of H_2O_2 or H_2O.

4.2 Potential energy profiles

A **potential energy profile** shows how the enthalpy of the system changes as reactants turn into products. The horizontal axis is the **reaction coordinate**, which represents the progress of the reaction (not time). Before reactants can form products, some bonds must start to break, so the energy rises to a maximum (the transition state) before it falls to the product level. The energy needed to reach this maximum is the **activation energy, E_a**: the minimum energy that colliding particles need in order to react.

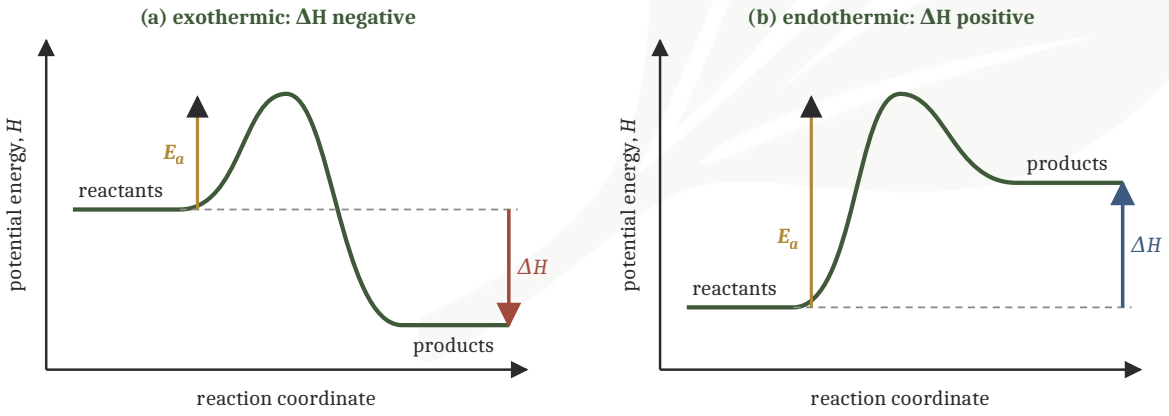


Figure 4 — Potential energy profiles. E_a is measured from the reactant level up to the peak. ΔH is measured from the reactant level to the product level, so its arrow points down for an exothermic reaction and up for an endothermic one.

How to earn every mark on a sketched profile

- **Axes:** y = "potential energy" (or "enthalpy, H"); x = "reaction coordinate" (or "progress of reaction"). Never "time".
- **Levels:** reactants and products labelled (with formulas if the equation is given), products **lower** for exothermic and **higher** for endothermic.
- **ΔH arrow:** from the reactant level to the product level, labelled ΔH with its sign or value.
- **E_a arrow:** from the reactant level to the top of the curve, labelled E_a.
- One smooth hump. Scale is not required unless values are given.

4.3 What ΔH cannot tell you

The sign of ΔH is a **guide** to the likely direction of change, not a complete one. Two limits are worth knowing.

1 · Some endothermic reactions are spontaneous. Dissolving ammonium nitrate happens on its own, and so does this strongly endothermic reaction:

$$SOCl_2(l) + FeCl_3{\cdot}6H_2O(s) \rightarrow FeCl_3(s) + 6SO_2(g) + 12HCl(g)$$

Endothermic changes like this can only happen when they produce a large increase in the dispersal (disorder) of matter and energy, here by releasing 18 mol of gas. This is the idea of **entropy**, developed in Reactivity 1.4.

2 · ΔH says nothing about rate. A mixture of methane and air is energetically unstable, yet it does not react until a match supplies the activation energy. Diamond is energetically unstable relative to graphite ($\Delta H^{\ominus} \approx -2$ kJ mol^{-1} for C(diamond) → C(graphite)), but the activation energy is so large that the change does not take place at an observable rate. We say diamond is **energetically unstable but kinetically stable**. Activation energy is developed in Reactivity 2.2.

Worked example — sketching and interpreting a profile

Q. Hydrogen peroxide decomposes: $2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$, ΔH = −196 kJ. The activation energy of the uncatalysed reaction is about 75 kJ mol^{-1}. (a) Sketch the energy profile. (b) Deduce which side is more stable. (c) Explain why a bottle of hydrogen peroxide can be stored for months. (d) Deduce the activation energy of the reverse reaction.

A. (a) Exothermic shape as in Figure 4(a): reactant level labelled $2H_2O_2(l)$, product level labelled $2H_2O(l) + O_2(g)$, drawn 196 kJ below the reactant level; peak drawn 75 kJ above the reactant level with the E_a arrow; ΔH = −196 kJ arrow pointing down from reactants to products.
(b) The products, water and oxygen, have lower enthalpy, so they are more stable.
(c) The activation energy is high enough that, at room temperature, very few collisions have enough energy to react, so the rate is very low even though the reaction is energetically favourable.
(d) The reverse reaction must climb from the product level to the same peak: E_a(reverse) = 75 + 196 = **271 kJ**.

Model answer: "Explain why the combustion of methane is exothermic, yet a mixture of methane and air does not ignite at room temperature." (3 marks)

The products, CO_2 and H_2O, have lower enthalpy than the reactants, so they are more stable and heat is released to the surroundings **[1]**. Bonds in the reactants must first be broken, so the reaction has a high activation energy **[1]**. At room temperature too few molecules have energy equal to or greater than E_a; a spark or flame supplies it **[1]**.

Examiner traps in 1.1.3

- Labelling the x-axis "time" loses the mark. The reaction coordinate is the progress of the reaction.
- Drawing the ΔH arrow from the **peak** to the products gives E_a(reverse), not ΔH.
- "The products are stable" is incomplete. Say the products are **more stable than the reactants** because they have **lower enthalpy**.
- A negative ΔH does not guarantee a reaction will happen at an observable rate.

5 Reactivity 1.1.4 — Standard enthalpy changes and thermochemical equations SL + HL

5.1 Standard conditions

The enthalpy change of a reaction is the heat transferred at **constant pressure**. Because its value depends on pressure, concentration and the physical states of the substances, values are compared under agreed standard conditions and given the symbol $\mathbf{\Delta H^{\ominus}}$ ("delta H standard").

Definition to learn — standard conditions

- a pressure of **100 kPa** (1.00×10^5 Pa)
- a concentration of **1 mol dm^{-3}** for all solutions
- all substances in their **standard states**: the pure, most stable form of the substance under standard conditions, e.g. $H_2O(l)$, $O_2(g)$, C(graphite), Na(s)

Temperature is **not** part of the definition, but values are usually quoted at **298 K (25 °C)**. Units of $\Delta H^{\ominus}$: **kJ mol^{-1}**.

5.2 Thermochemical equations

A thermochemical equation is a balanced equation, with state symbols, followed by its enthalpy change:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \quad \Delta H^{\ominus} = -891 \text{ kJ mol}^{-1}$$

This is shorthand for: one mole of methane gas reacts with two moles of oxygen gas to form one mole of carbon dioxide gas and two moles of liquid water, and 891 kJ of heat is **released**. The "per mol" means per mole of the reaction **as written**. Similarly, photosynthesis absorbs energy:

$$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g) \quad \Delta H^{\ominus} = +2803 \text{ kJ mol}^{-1}$$

If you...	...then ΔH	Example (from the methane equation)
multiply the equation by a factor	is multiplied by the same factor	$2CH_4 + 4O_2 \rightarrow 2CO_2 + 4H_2O$: $\Delta H = -1782$ kJ
reverse the equation	changes sign, same size	$CO_2 + 2H_2O \rightarrow CH_4 + 2O_2$: $\Delta H = +891$ kJ
change a state symbol	changes value	forming $2H_2O(g)$ instead of $2H_2O(l)$ is less exothermic, because condensing water releases extra heat

5.3 Two named enthalpy changes you need now

Definitions to learn

Standard enthalpy of combustion, $\Delta H_c^{\ominus}$: the enthalpy change when **one mole** of a substance is burned completely in oxygen under standard conditions. It is always negative.

Enthalpy of neutralization: the enthalpy change when **one mole of water** is formed by the reaction of an acid with a base. For any strong acid with any strong base it is about −57 kJ mol^{-1}, because the reaction is always $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$.

Because ΔH_c is defined per mole of fuel, its equation must have a coefficient of exactly 1 in front of the fuel, even if that means fractions elsewhere: $CH_3OH(l) + 1\frac{1}{2}O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$, $\Delta H_c^{\ominus} = -726$ kJ mol^{-1}.

Worked example — using a thermochemical equation

Q. Use $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$, $\Delta H^{\ominus} = -891$ kJ mol^{-1}. (a) Calculate the heat released when 4.00 g of methane burns completely. (b) Calculate the minimum mass of methane needed to heat 2.00 kg of water by 50.0 K.

A. (a) $n(CH_4) = 4.00 / 16.05 = 0.249$ mol; heat released = $0.249 \times 891 =$ **222 kJ**.
(b) $Q = mc\Delta T = 2000\text{ g} \times 4.18\text{ J g}^{-1}\text{ K}^{-1} \times 50.0\text{ K} = 418\,000$ J = 418 kJ.
$n(CH_4) = 418 / 891 = 0.469$ mol; $m = 0.469 \times 16.05 =$ **7.53 g** (assuming all the heat reaches the water).

Examiner traps in 1.1.4 definitions

- Missing “one mole” or “completely” from the ΔH_c definition, or writing a combustion equation for 2 mol of fuel.
- Defining enthalpy of neutralization per mole of acid: it is per mole of **water formed**. For H_2SO_4, one mole of acid makes two moles of water.
- Quoting 25 °C as part of the definition of standard conditions; it is the usual temperature, not part of the definition.
- Leaving out state symbols in a thermochemical equation.

6 Reactivity 1.1.4 — Calculating heat with Q = mcΔT SL + HL

6.1 Specific heat capacity

If the same amount of heat is added to two objects, the temperature change is not necessarily the same. The added energy is shared among all the particles, so an object with fewer particles experiences a larger rise in temperature. The temperature change, ΔT, therefore depends on the **mass** of the object, the **heat added**, Q, and the **nature of the substance**. The third factor is captured by its specific heat capacity.

> **Definition to learn**
>
> The **specific heat capacity, c,** of a substance is the heat needed to raise the temperature of a unit mass (1 g) of the substance by 1 K. Units: **J g^{-1} K^{-1}** (identical to J g^{-1} °C^{-1}).

$$Q = m \times c \times \Delta T$$

Q in J · m in g · c in J g^{-1} K^{-1} · ΔT in K (or °C) · for water, c = 4.18 J g^{-1} K^{-1} (data booklet)

Substance	water	ethanol	aluminium	iron	copper	mercury	gold
c / J g^{-1} K^{-1}	4.18	2.44	0.897	0.449	0.385	0.138	0.129

The lower the specific heat capacity, the larger the temperature change for a given amount of heat. Metals heat up and cool down quickly; water does not.

> **Why water's specific heat capacity is so high**
>
> Water molecules are held together by hydrogen bonds. Some of the heat supplied is used to stretch and break these bonds rather than to increase the kinetic energy of the molecules, so more energy is needed for each 1 K rise. Also, 1 g of water contains a relatively large number of small molecules. This is why water is a good coolant, why coastal climates are mild, and why it is the liquid of choice in calorimeters.

For a whole object such as a copper calorimeter, the **heat capacity** C = mc (unit J K^{-1}) gives the heat needed to raise the object by 1 K.

Worked examples — four uses of Q = mcΔT

Question	Working	Answer
Heat needed to warm a 25.0 g aluminium block (c = 0.897) from 21.0 °C to 95.0 °C	Q = 25.0 × 0.897 × (95.0 − 21.0) = 25.0 × 0.897 × 74.0 = 1659 J	**1.66 kJ**
Heat released when 10.0 g of copper (c = 0.385) cools from 85.0 °C to 25.0 °C	Q = 10.0 × 0.385 × 60.0 = 231 J. The copper *loses* this heat: ΔT for the copper is −60.0 K.	**231 J released**
Final temperature when 150.0 J is added to 50.0 g of mercury (c = 0.138) at 22.0 °C	ΔT = Q / mc = 150.0 / (50.0 × 0.138) = 21.7 K; T = 22.0 + 21.7	**43.7 °C**
2.50 kJ raises 200.0 g of a liquid by 5.12 K. Find c and identify the liquid.	c = Q / mΔT = 2500 / (200.0 × 5.12) = 2.44 J g^{-1} K^{-1}, matching ethanol	**2.44; ethanol**

Comparing water and copper: equal masses receiving equal heat give ΔT(Cu) / ΔT(H_2O) = 4.18 / 0.385 = 10.9, so the copper's temperature rises almost **11 times** as much.

6.2 From Q to ΔH

In a calorimetry experiment the heat measured, Q, is the heat **gained or lost by the water**. The reaction's enthalpy change has the opposite sign, and it must be expressed per mole:

$$\Delta H = -Q / n$$

n = amount (mol) of the substance named in the definition, or of the limiting reactant

Method — the four steps for every calorimetry calculation

1. **Q = mcΔT**, using the mass of the **water or solution** that changed temperature (never the mass of fuel or solid).
2. Convert to kJ: divide by 1000.
3. Find **n**: n = m/M for a fuel or solid, n = c × V/1000 for a solution. Use the **limiting** reactant, or the moles of water formed for neutralization.
4. **ΔH = −Q/n**. Temperature rose → ΔH negative. Temperature fell → ΔH positive. Give 3 significant figures and units kJ mol^{-1}.

Water's temperature	Heat transfer	Q (for the water)	ΔH (for the reaction)
rises	system → water	positive	**negative** (exothermic)
falls	water → system	negative (Q = mcΔT with ΔT < 0)	**positive** (endothermic)

7 Reactivity 1.1.4 — Calorimetry I: enthalpy of combustion SL + HL

7.1 The method

The enthalpy of combustion of a liquid fuel such as an alcohol can be found with simple apparatus. The flame heats a known mass of water in a metal calorimeter; copper is often used because it is a good conductor, so heat passes quickly from the flame to the water.

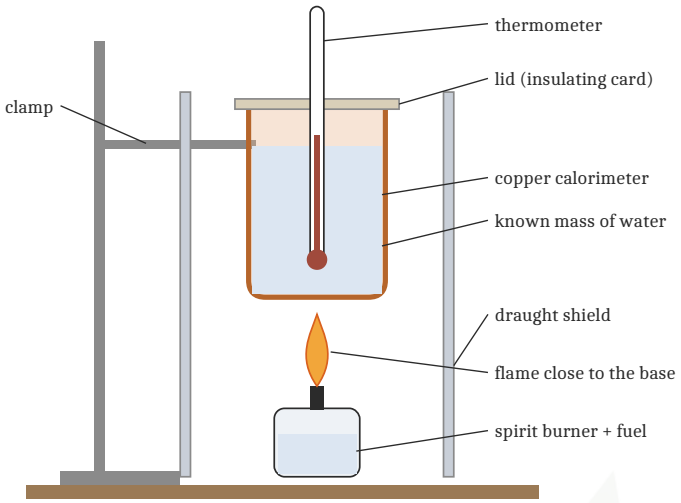


Figure 5 — Measuring the enthalpy of combustion of a liquid fuel. The draught shield and lid reduce heat loss; the flame is placed close to the base of the calorimeter.

1. Weigh the spirit burner containing the fuel, with its cap on.
2. Measure a known volume (mass) of water into the copper calorimeter and record its initial temperature.
3. Light the burner and heat the water, stirring, until the temperature has risen by about 10–20 K.
4. Put the cap on to extinguish the flame, record the **highest** temperature reached, and reweigh the burner straight away.

Worked example 1 — methanol (water only)

Mass of methanol burned / g	0.640
Mass of water in calorimeter / g	150.0
Temperature rise of water / K	14.2

Q. Calculate the enthalpy of combustion of methanol, M = 32.05 g mol^{-1}, and the percentage error compared with the data booklet value of −726 kJ mol^{-1}.

$Q = mc\Delta T = 150.0 \times 4.18 \times 14.2 = 8903\ \text{J} = 8.90\ \text{kJ}$

$n(CH_3OH) = 0.640 / 32.05 = 0.01997\ \text{mol}$

$\Delta H_c = -Q/n = -8.90 / 0.01997 =$ **$-446\ \text{kJ mol}^{-1}$**

percentage error $= |-446 - (-726)| / |-726| \times 100 = 280 / 726 \times 100 =$ **38.6 %**

The negative sign shows the reaction is exothermic: the water gained the heat that the system lost. The experimental value is much **less exothermic** than the literature value, which is typical of this apparatus (section 7.3).

Worked example 2 — including the heat absorbed by the calorimeter

Q. Burning 0.300 g of propan-1-ol (M = 60.11 g mol^{-1}) raised the temperature of 100.0 g of water in an 80.0 g copper calorimeter (c = 0.385 J g^{-1} K^{-1}) by 18.5 K. Calculate ΔH_c, first ignoring and then including the calorimeter.

Heat absorbed by water = 100.0 × 4.18 × 18.5 = 7733 J

Heat absorbed by copper = 80.0 × 0.385 × 18.5 = 570 J

n(C_3H_7OH) = 0.300 / 60.11 = 4.991 × 10^{-3} mol

Water only: ΔH_c = −7.733 / 4.991 × 10^{-3} = −1.55 × 10^{3} kJ mol^{-1}

Water + calorimeter: Q = 7733 + 570 = 8303 J; ΔH_c = −8.303 / 4.991 × 10^{-3} = **−1.66 × 10^{3} kJ mol^{-1}**

Including the calorimeter brings the value closer to the data booklet value (−2021 kJ mol^{-1}), but a large gap remains, mainly because of heat lost to the air.

7.2 Percentage error

percentage error = |experimental value − accepted value| / |accepted value| × 100 %

7.3 Why the experimental value is less exothermic

The data booklet value is measured in a sealed bomb calorimeter, under standard conditions, with complete combustion and water formed as a liquid. A spirit burner heating an open calorimeter departs from all of these. Learn each error together with the **direction** of its effect and a **specific** improvement.

Source of error	Effect on the result	Improvement
Heat lost to the air, the clamp and the stand	less heat reaches the water, so Q and ΔT are too small: value less exothermic	draught shield, lid, insulate the calorimeter, flame close to the base
Incomplete combustion (yellow, sooty flame)	less energy released per mole than complete combustion: less exothermic	increase the oxygen supply; evidence is soot on the calorimeter base
Heat absorbed by the calorimeter not counted	Q underestimated: less exothermic	add m c ΔT for the calorimeter (worked example 2)
Fuel evaporates from the wick	mass loss is too large, so n is too large: less exothermic	keep the cap on whenever the burner is not lit; reweigh immediately
Non-standard conditions; some water leaves as vapour	value differs from $\Delta H^{\ominus}$; vapour carries energy away	a bomb calorimeter is needed for accurate values

Avoid trivial answers

- "Human error", "inaccurate measurement" and "repeat the experiment" do not score. Name the error, say which way it moves the result, and link it to an improvement.
- Heat loss is a **systematic** error: it always pushes the result in the same direction, so repeating the experiment will not remove it. Repeats only reduce **random** errors, such as reading the thermometer.
- Record qualitative observations too: soot and a yellow flame are evidence of incomplete combustion.

Model answer: "A student obtained −1050 kJ mol^{-1} for the enthalpy of combustion of ethanol. The data booklet value is −1367 kJ mol^{-1}. Suggest two reasons for the difference and one improvement." (3 marks)

Heat was lost to the surroundings, so less heat was absorbed by the water and the calculated value is less exothermic **[1]**. Combustion of the ethanol was incomplete (soot formed), so less energy was released per mole **[1]**. Use a draught shield and a lid and insulate the calorimeter to reduce heat loss **[1]**.

8 Reactivity 1.1.4 — Calorimetry II: reactions in solution SL + HL

8.1 The coffee-cup calorimeter and its assumptions

For reactions in solution, the reaction is carried out in an expanded polystyrene cup. Polystyrene is a very good insulator with a very low heat capacity, so nearly all the heat transferred passes between the reaction and the water in the solution, which acts as the surroundings.

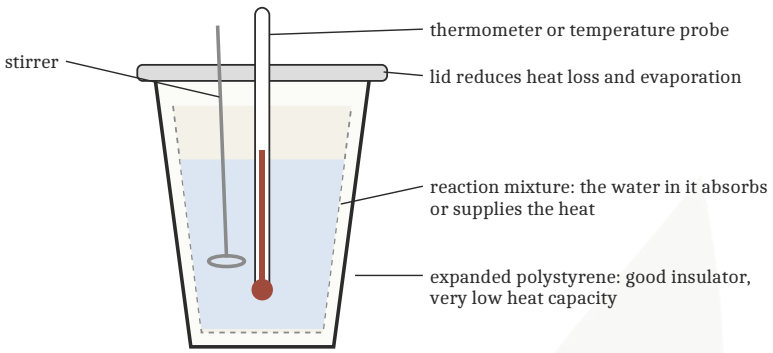


Figure 6 — A simple solution calorimeter. The water in the solution absorbs heat from an exothermic reaction or supplies heat to an endothermic one.

The standard assumptions (state them when asked)

1. The solutions are dilute, so their density is that of water, **1.00 g cm^{-3}**: mass in g = **total** volume in cm^3.
2. The specific heat capacity of the solution is that of water, **4.18 J g^{-1} K^{-1}**.
3. All the heat is transferred between the reaction and the water; the cup absorbs a negligible amount.
4. No heat is lost to the surroundings (or the loss is corrected for, section 8.2).

Worked example 3 — enthalpy of neutralization

Q. 50.0 cm^3 of 1.00 mol dm^{-3} HCl(aq) and 50.0 cm^3 of 1.00 mol dm^{-3} NaOH(aq), both at 21.4 °C, were mixed in a polystyrene cup. The highest temperature was 28.1 °C. Calculate the enthalpy of neutralization.

total mass of solution = 50.0 + 50.0 = 100.0 g; ΔT = 28.1 − 21.4 = 6.7 K

$Q = 100.0 \times 4.18 \times 6.7 = 2801$ J = 2.80 kJ

$n(H_2O) = n(HCl) = n(NaOH) = 1.00 \times 50.0/1000 = 0.0500$ mol

$\Delta H = -2.80 / 0.0500 =$ **−56.0 kJ mol^{-1}**

Worked example 4 — the limiting reactant and a diprotic acid

Q. 30.0 cm^3 of 1.00 mol dm^{-3} $H_2SO_4(aq)$ was mixed with 50.0 cm^3 of 1.00 mol dm^{-3} NaOH(aq). The temperature rose by 8.3 K. Calculate the enthalpy of neutralization.

$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

$n(NaOH) = 0.0500$ mol; $n(H_2SO_4) = 0.0300$ mol, which would need 0.0600 mol NaOH → **NaOH is limiting**

$n(H_2O$ formed$) = n(NaOH) = 0.0500$ mol

$Q = (30.0 + 50.0) \times 4.18 \times 8.3 = 2776$ J = 2.78 kJ

$\Delta H = -2.78 / 0.0500 =$ **-55.5 kJ mol^{-1}** (per mole of water formed)

Dividing by the moles of acid (0.0300) would give a wrong answer. The definition is per mole of water, and only the limiting reactant controls how much water forms.

Worked example 5 — an endothermic dissolving

Q. 4.00 g of ammonium nitrate (M = 80.06 g mol^{-1}) was dissolved in 50.0 g of water. The temperature fell from 22.0 °C to 16.3 °C. Calculate the enthalpy change of solution.

$\Delta T = 16.3 - 22.0 = -5.7$ K; heat lost by the water $= 50.0 \times 4.18 \times 5.7 = 1191$ J = 1.19 kJ

$n(NH_4NO_3) = 4.00 / 80.06 = 0.0500$ mol

The water lost heat to the system, so ΔH is positive: $\Delta H = +1.19 / 0.0500 =$ **+23.8 kJ mol^{-1}**

Here m is the mass of the **water** (the solid's mass is often ignored; if a question uses the mass of the whole solution, 54.0 g, follow it). The literature value is about +25.7 kJ mol^{-1}; heat flowing *in* from the room makes the measured ΔT too small.

8.2 Correcting for heat loss: extrapolating a cooling curve

Heat starts to escape as soon as the solution is warmer than the room, so the highest recorded temperature is always lower than the temperature an ideal, perfectly insulated system would reach. We correct for this by recording the temperature at regular intervals before and after the reaction, then **extrapolating the cooling section back to the time of mixing**. This assumes the solution cools at the same rate throughout.

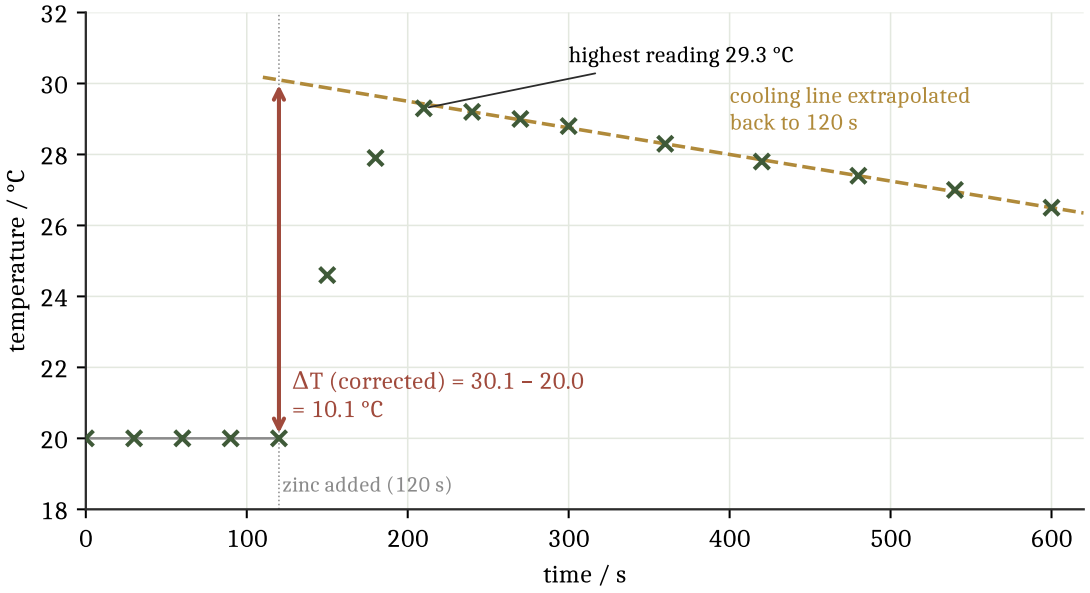


Figure 7 — Temperature against time for zinc powder added to copper(II) sulfate solution at 120 s. The cooling line, extrapolated back to 120 s, gives the temperature the solution would have reached with no heat loss.

Worked example 6 — displacement with a corrected ΔT

Q. 50.0 cm^3 of 0.200 mol dm^{-3} $CuSO_4(aq)$ was placed in a polystyrene cup and 1.00 g of zinc powder was added at 120 s (Figure 7). $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$. Calculate ΔH, and compare it with the value from the highest recorded temperature.

Limiting reactant: $n(Cu^{2+}) = 0.200 \times 50.0/1000 = 0.0100$ mol; $n(Zn) = 1.00 / 65.38 = 0.0153$ mol → **Cu^{2+} limiting** (zinc in excess)

Corrected $\Delta T = 30.1 - 20.0 = 10.1$ K; $Q = 50.0 \times 4.18 \times 10.1 = 2111$ J = 2.11 kJ

$\Delta H = -2.11 / 0.0100 =$ **-211 kJ mol^{-1}**

Uncorrected: $\Delta T = 29.3 - 20.0 = 9.3$ K gives −194 kJ mol^{-1}, further from the literature value of about −218 kJ mol^{-1}.

8.3 The mixing (thermometric) method

If the total volume is kept constant and the proportions of two reacting solutions are varied, the largest temperature rise occurs when the reactants are in their **stoichiometric ratio**: neither is in excess, so the most product forms in the smallest volume. Two best-fit lines are drawn and their intersection gives both the volumes that react exactly and the maximum ΔT.

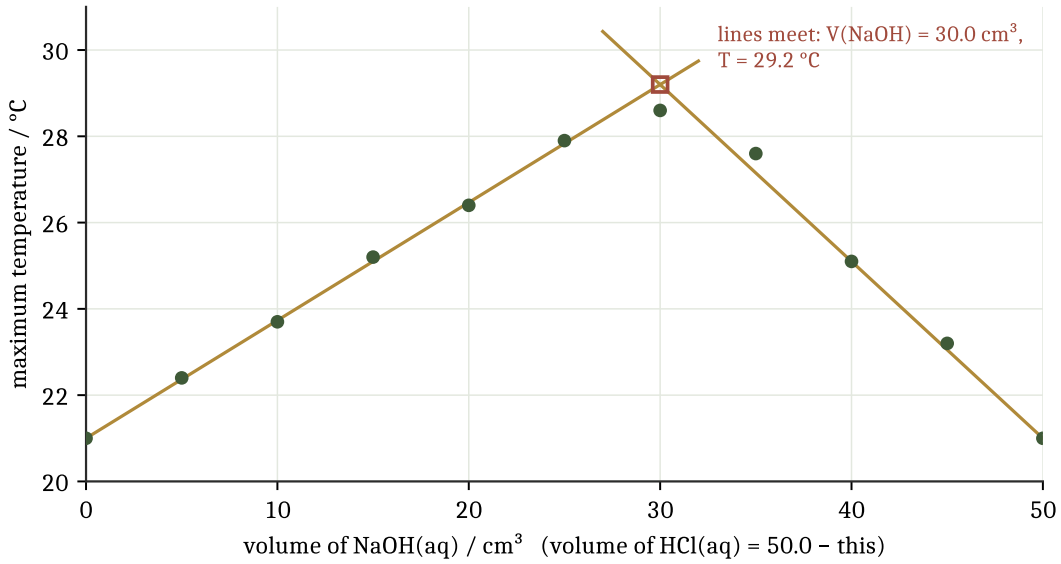


Figure 8 — Maximum temperature against volume of 1.00 mol dm^{-3} NaOH(aq) added to HCl(aq) of unknown concentration, total volume 50.0 cm^3, initial temperature 21.0 °C. The points near the peak lie below the lines because of heat loss, so the intersection of the lines is used.

Worked example 7 — finding a concentration and ΔH from a mixing graph

Lines meet at V(NaOH) = 30.0 cm^3, so V(HCl) = 50.0 − 30.0 = 20.0 cm^3; ΔT = 29.2 − 21.0 = 8.2 K

n(NaOH) = 1.00 × 30.0/1000 = 0.0300 mol = n(HCl) (1 : 1) → [HCl] = 0.0300 / 0.0200 = **1.50 mol dm^{-3}**

Q = 50.0 × 4.18 × 8.2 = 1714 J = 1.71 kJ; n(H_2O) = 0.0300 mol

ΔH = −1.71 / 0.0300 = **−57.1 kJ mol^{-1}**

Model answer: "Explain why the temperature change was found by extrapolation rather than from the highest temperature recorded." (2 marks)

Heat is lost to the surroundings while the reaction is still taking place, so the highest recorded temperature is lower than it would be with no heat loss **[1]**. Extrapolating the cooling line back to the time of mixing estimates the temperature that would have been reached if the reaction were instantaneous with no heat loss, assuming the rate of cooling is constant **[1]**.

Examiner traps in solution calorimetry

- Using the volume of only **one** solution for m. After mixing, the whole solution changes temperature: use the total volume.
- Using the mass of the solid or the fuel as m in Q = mcΔT.
- Dividing by the moles of the reactant in **excess**. Always check which reactant is limiting first.
- Forgetting the negative sign for an exothermic reaction, or leaving Q in J while quoting kJ mol^{-1}.
- Quoting too many significant figures. A ΔT of 6.7 K (2 s.f.) does not justify an answer of −56.012 kJ mol^{-1}.

9 Key points to memorise

Everything below is either wording that examiners look for or knowledge you must carry into the exam yourself.

A Definitions — near-verbatim

- **Heat:** energy transferred because of a temperature difference. **Temperature:** a measure of the average kinetic energy of the particles.
- **System:** the part being studied (the reacting chemicals). **Surroundings:** everything else. Total energy is conserved.
- **Exothermic:** heat transferred from the system to the surroundings; ΔH negative; temperature of surroundings rises. **Endothermic:** the reverse; ΔH positive; temperature falls.
- **Enthalpy, H:** the chemical potential energy stored in a system. **ΔH** = H(products) − H(reactants) = heat transferred at constant pressure.
- **Standard conditions:** 100 kPa, 1 mol dm^{-3} solutions, all substances in standard states (usually at 298 K).
- **Specific heat capacity:** heat needed to raise the temperature of 1 g of a substance by 1 K.
- **$\Delta H_c^{\ominus}$:** enthalpy change when one mole of a substance is completely burned in oxygen under standard conditions. **Enthalpy of neutralization:** per mole of water formed.

B Equations and sign rules

- $Q = mc\Delta T$ (J); $\Delta H = -Q/n$ (kJ mol^{-1} after dividing Q by 1000). c(water) = 4.18 J g^{-1} K^{-1}.
- Temperature rises → ΔH negative. Temperature falls → ΔH positive.
- Multiply an equation → multiply ΔH. Reverse an equation → change the sign of ΔH.
- Lower enthalpy = more stable. Exothermic: products more stable. Endothermic: reactants more stable.
- Energy profile: y = potential energy, x = reaction coordinate; E_a from reactants to peak; ΔH from reactants to products.

C Practical knowledge

- Solution assumptions: density 1.00 g cm^{-3}, c of solution = c of water, cup absorbs negligible heat, no heat loss.
- Combustion errors: heat loss, incomplete combustion, heat absorbed by calorimeter, fuel evaporation, non-standard conditions. All make the value **less exothermic**.
- Heat loss is systematic: correct it by insulation, a lid, a draught shield, or by extrapolating a cooling curve. Repeats do not remove it.
- Percentage error = |experimental − accepted| / |accepted| × 100 %.

10 Check yourself

Use the data booklet where needed: c(water) = 4.18 J g^{-1} K^{-1}. Marks are shown in brackets. Attempt the questions under timed conditions: roughly 1.2 minutes per mark. Full answers with mark-scheme wording are in section 11.

Part A — Multiple choice SL + HL

1. Which statement about heat and temperature is correct?

A Heat and temperature are both measured in joules.
B Temperature is a measure of the total kinetic energy of the particles in an object.
C Heat is energy transferred because of a temperature difference.
D An object at a higher temperature always contains more heat than one at a lower temperature.

2. The temperature of some water rises from 291 K to 306 K. What is the temperature rise?

A 15 °C **B** 288 °C **C** 306 °C **D** 579 °C

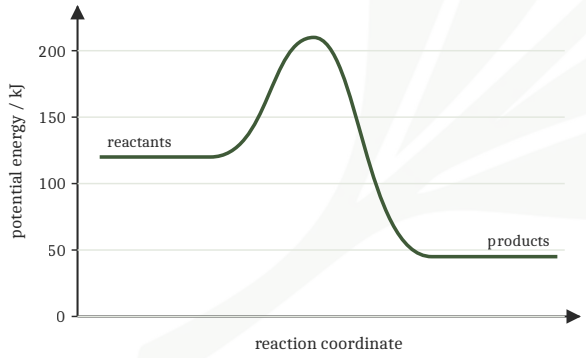
3. When ammonium chloride dissolves in water, the beaker feels cold. Which row is correct?

	Type of process	Sign of ΔH	Direction of heat transfer
A	exothermic	negative	system → surroundings
B	endothermic	positive	surroundings → system
C	endothermic	negative	surroundings → system
D	exothermic	positive	system → surroundings

4. Which process is exothermic?

A melting ice **B** $CaCO_3 \rightarrow CaO + CO_2$ **C** evaporating ethanol **D** condensing steam

5. The potential energy profile for a reaction is shown. Which values are correct?



A E_a = 90 kJ, ΔH = −75 kJ
B E_a = 210 kJ, ΔH = −75 kJ
C E_a = 90 kJ, ΔH = +75 kJ
D E_a = 165 kJ, ΔH = −45 kJ

6. Which describes standard conditions for enthalpy changes?

A 100 kPa, 1 mol dm^{-3} for solutions, all substances in their standard states
B 101.3 kPa and 273 K only
C 100 kPa and 298 K only
D 1 mol dm^{-3} for solutions and 273 K, at any pressure

7. 1.00 kJ of heat is added to 50.0 g of each substance. Which has the largest temperature rise?

A water (4.18) **B** ethanol (2.44) **C** aluminium (0.897) **D** silver (0.237)

specific heat capacities in J g^{-1} K^{-1}

8. How much heat is needed to raise the temperature of 20.0 g of water from 18.0 °C to 43.0 °C?

A 0.418 kJ **B** 2.09 kJ **C** 3.59 kJ **D** 25.0 kJ

9. Mixing 50.0 cm^3 of 1.00 mol dm^{-3} HCl with 50.0 cm^3 of 1.00 mol dm^{-3} NaOH gives a temperature rise of 6.8 K. What is the rise when 100.0 cm^3 of each of the same solutions are mixed?

A 3.4 K **B** 6.8 K **C** 13.6 K **D** 27.2 K

10. Burning 0.460 g of ethanol (M = 46.08 g mol^{-1}) raised the temperature of 100.0 g of water by 12.0 K. Which expression gives ΔH_c in kJ mol^{-1}?

A $-(100.0 \times 4.18 \times 12.0) \div (0.460 / 46.08)$
B $-(100.0 \times 4.18 \times 12.0) \div (1000 \times 0.460 / 46.08)$
C $-(0.460 \times 4.18 \times 12.0) \div (1000 \times 100.0 / 46.08)$
D $-(100.0 \times 4.18 \times 285) \div (1000 \times 0.460 / 46.08)$

11. Which would give an experimental ΔH_c that is **less exothermic** than the data booklet value?

I. heat lost from the sides of the calorimeter
II. fuel evaporating from the wick before the burner is reweighed
III. recording the mass of water as larger than it really was

A I and II only **B** I and III only **C** II and III only **D** I, II and III

12. $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ $\Delta H^\ominus = -572$ kJ mol^{-1}. How much heat is released when 1.00 g of hydrogen burns completely?

A 142 kJ **B** 283 kJ **C** 286 kJ **D** 572 kJ

13. For C(diamond) → C(graphite), $\Delta H^\ominus \approx -2$ kJ mol^{-1}. Why does diamond not turn into graphite at room temperature?

A Graphite is less stable than diamond.
B The reaction is endothermic.
C The activation energy is very high.
D ΔH is too small for the reaction to occur.

14. Why is an expanded polystyrene cup used for reactions in solution rather than a glass beaker?

A It does not react with acids or bases.
B It is a good thermal insulator with a very low heat capacity.
C It has a high specific heat capacity, so it absorbs the heat released.
D It conducts heat quickly from the solution to the thermometer.

Part B — Structured questions SL + HL

15. (a) Distinguish between heat and temperature. **[2]**

(b) A kettle of water and a swimming pool are both at 20 °C. Explain why far more heat is needed to raise the temperature of the pool by 1 K. **[1]**

(c) State, with a reason, which contains particles with the greater average kinetic energy. **[1]**

16. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ $\Delta H^{\ominus} = +178$ kJ mol^{-1}

(a) Sketch a labelled potential energy profile for this reaction, showing ΔH and E_a. [3]

(b) Deduce, with a reason, whether the reactant or the products are more stable. [1]

(c) Calculate the heat needed to decompose 25.0 g of calcium carbonate (M = 100.09 g mol^{-1}). [1]

17. The standard enthalpy of combustion of propane, C_3H_8, is −2219 kJ mol^{-1}.

(a) Write the thermochemical equation for the standard enthalpy of combustion of propane. [1]

(b) Calculate the heat released when 11.0 g of propane burns completely. [2]

(c) Calculate the mass of water, in kg, that this heat could warm from 20.0 °C to 100.0 °C, assuming no heat is lost. [2]

18. A student burned butan-1-ol (M = 74.14 g mol^{-1}) under a copper calorimeter containing water.

Initial mass of burner and butan-1-ol / g	52.418
Final mass of burner and butan-1-ol / g	52.106
Mass of water / g	200.0
Initial temperature of water / °C	19.6
Highest temperature of water / °C	31.4

(a) Calculate the heat absorbed by the water, in kJ. [1]

(b) Calculate the amount, in mol, of butan-1-ol burned. [1]

(c) Determine the enthalpy of combustion of butan-1-ol, in kJ mol^{-1}. [1]

(d) The data booklet value is −2676 kJ mol^{-1}. Calculate the percentage error. [1]

(e) Suggest two reasons for the difference, stating the effect of each on the result. [2]

(f) Suggest one specific improvement linked to one of your reasons. [1]

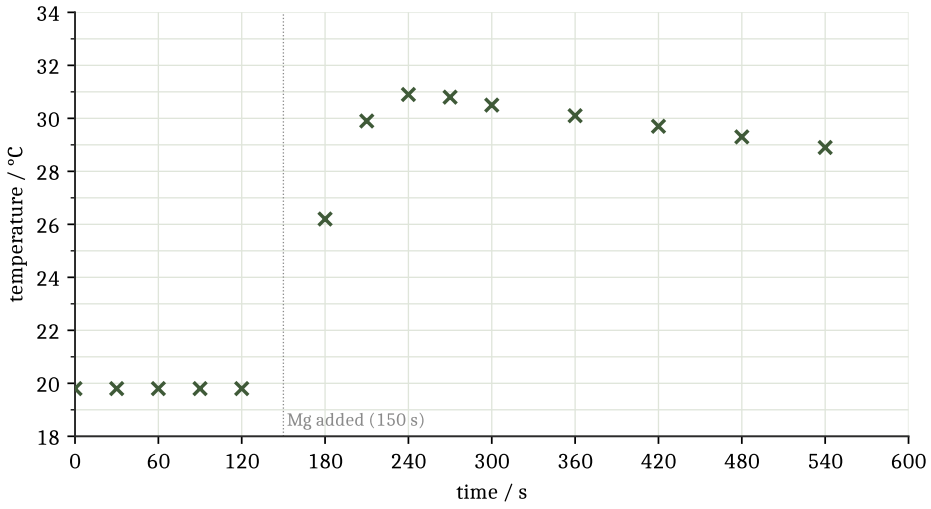
(g) Black soot formed on the base of the calorimeter. State what this shows. [1]

19. 25.0 cm^3 of 2.00 mol dm^{-3} $HNO_3(aq)$ was mixed with 50.0 cm^3 of 1.00 mol dm^{-3} KOH(aq). The temperature rose by 9.0 K.

(a) Calculate the enthalpy of neutralization. [3]

(b) ★ Predict the temperature rise when 25.0 cm^3 of the same $HNO_3(aq)$ is mixed with 25.0 cm^3 of the same KOH(aq). Explain your reasoning. [2]

20. Magnesium powder (in excess) was added at 150 s to 50.0 cm^3 of 0.100 mol dm^{-3} $CuSO_4(aq)$ in a polystyrene cup. The temperature was recorded every 30 s.



(a) Write the ionic equation for the reaction. [1]

(b) By extrapolating the graph, determine the temperature change that would have occurred with no heat loss. [2]

(c) Calculate the heat produced, in kJ. [1]

(d) Calculate the enthalpy change of the reaction in kJ mol^{-1}. [2]

(e) Explain why the value in (b) is larger than the change found from the highest reading. [1]

21. ★ In a mixing experiment, different volumes of 1.20 mol dm^{-3} NaOH(aq) and $H_2SO_4(aq)$ of unknown concentration were mixed, keeping the total volume at 60.0 cm^3. Both solutions started at 22.0 °C. The two best-fit lines met at V(NaOH) = 40.0 cm^3 and a temperature of 32.6 °C.

(a) State the volume of $H_2SO_4(aq)$ at the intersection and explain why the temperature rise is greatest at this point. [2]

(b) Calculate the concentration of the sulfuric acid. [2]

(c) Calculate the enthalpy of neutralization per mole of water formed. [2]

22. ★ A 25.00 g iron bolt (c = 0.449 J g^{-1} K^{-1}) was held in a Bunsen flame for several minutes, then dropped quickly into 150.0 g of water at 20.0 °C in a polystyrene cup. The final temperature was 38.4 °C.

(a) Calculate the heat gained by the water. [1]

(b) Determine the temperature of the flame, stating one assumption. [2]

(c) Suggest whether the true flame temperature is higher or lower than your answer, with a reason. [1]

23. 5.05 g of potassium nitrate (M = 101.11 g mol^{-1}) was dissolved in 50.0 g of water. The temperature fell from 21.4 °C to 14.6 °C.

(a) Calculate the enthalpy change of solution of potassium nitrate. [3]

(b) State whether the solid or the solution has the greater enthalpy. [1]

24. A student measured the enthalpy of neutralization of HCl by NaOH as −45 kJ mol^{-1} (accepted value −57 kJ mol^{-1}) and concluded: "The difference was due to human error, so the experiment should be repeated more carefully." Evaluate this conclusion and suggest two specific improvements. [3]

Part A: 14 marks · Part B: 51 marks · Total: 65 marks

11 Answers

Part A

1. C. Heat is a process of energy transfer; temperature measures the average (not total) kinetic energy, and objects "contain" energy, not heat.

2. A. 306 − 291 = 15 K, and a change of 1 K equals a change of 1 °C.

3. B. The beaker feels cold because heat flows from the surroundings into the system: endothermic, ΔH positive.

4. D. Condensing forms intermolecular forces and releases heat. A, B and C all absorb heat.

5. A. E_a = 210 − 120 = 90 kJ; ΔH = 45 − 120 = −75 kJ.

6. A. Temperature is not part of the definition (298 K is usual); 101.3 kPa and 273 K describe older "STP" conventions for gases.

7. D. $\Delta T = Q/mc$, so the smallest c gives the largest rise.

8. B. Q = 20.0 × 4.18 × 25.0 = 2090 J = 2.09 kJ. (C uses 43.0 instead of ΔT.)

9. B. Twice the moles react, releasing twice the heat, but into twice the mass of solution, so ΔT is unchanged.

10. B. Q in J must be divided by 1000 to give kJ; n = 0.460/46.08. A gives J mol^{-1}; D wrongly adds 273 to ΔT.

11. A. I lowers Q; II makes n too large. III makes Q too large, giving a *more* exothermic value.

12. A. $n(H_2)$ = 1.00/2.02 = 0.495 mol; 572 kJ is for 2 mol H_2, so heat = 0.495 × 286 = 142 kJ.

13. C. Graphite is more stable, but the very high activation energy means the rate at room temperature is negligible.

14. B. Low heat capacity means the cup itself absorbs very little heat; good insulation reduces heat exchange with the room.

Part B

15. (a) Heat is energy transferred because of a temperature difference (J) **[1]**; temperature is a measure of the average kinetic energy of the particles (K) **[1]**.
(b) The pool contains far more water molecules, and each must gain the same increase in kinetic energy, so the total energy needed is much greater **[1]**.
(c) Neither: they are at the same temperature, so the average kinetic energy of their particles is the same **[1]**.

16. (a) Axes labelled potential energy (y) and reaction coordinate (x) **[1]**; product level $CaO(s) + CO_2(g)$ above reactant level $CaCO_3(s)$, with ΔH = +178 kJ arrow pointing up from reactants to products **[1]**; single hump with E_a arrow from reactant level to the peak **[1]**.
(b) $CaCO_3$ is more stable, because it has lower enthalpy than the products **[1]**.
(c) n = 25.0/100.09 = 0.250 mol; heat = 0.250 × 178 = **44.5 kJ** (absorbed) **[1]**.

17. (a) $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$, $\Delta H^{\ominus} = -2219$ kJ mol^{-1} (states and 1 mol propane required) **[1]**.
(b) n = 11.0/44.11 = 0.249 mol **[1]**; heat = 0.249 × 2219 = **553 kJ [1]**.
(c) $m = Q/c\Delta T$ = 553 000 / (4.18 × 80.0) **[1]** = 1.65 × 10^3 g = **1.65 kg [1]**.

18. (a) Q = 200.0 × 4.18 × 11.8 = 9865 J = **9.86 kJ [1]**.
(b) mass burned = 52.418 − 52.106 = 0.312 g; n = 0.312/74.14 = **4.21×10^{-3} mol [1]**.
(c) ΔH_c = −9.86 / 4.21×10^{-3} = **-2.34×10^{3} kJ mol^{-1}** (negative sign required) **[1]**.
(d) |−2344 − (−2676)| / 2676 × 100 = **12.4 %** (accept 12–13 %) **[1]**.
(e) Any two, each with its effect **[2]**: heat lost to the surroundings, so less heat reaches the water and Q is too small; incomplete combustion, so less energy is released per mole; heat absorbed by the copper calorimeter not included, so Q is underestimated; butan-1-ol evaporated from the wick, so the mass burned (and n) is overestimated. Each makes the value less exothermic.
(f) For example: use a draught shield and lid / insulate the calorimeter to reduce heat loss; or include m c ΔT for the copper; or keep the cap on the burner and reweigh immediately **[1]**.
(g) Combustion was incomplete (carbon formed because the oxygen supply was insufficient) **[1]**.

19. (a) n(HNO_3) = 0.0500 mol = n(KOH) = 0.0500 mol, so n(H_2O) = 0.0500 mol **[1]**; Q = 75.0 × 4.18 × 9.0 = 2822 J **[1]**; ΔH = −2.82/0.0500 = **−56.4 kJ mol^{-1} [1]**.
(b) KOH is now limiting: only 0.0250 mol of water forms, half the heat (1.41 kJ) **[1]**, released into 50.0 g of solution instead of 75.0 g: ΔT = 1411 / (50.0 × 4.18) = **6.8 K** (= 9.0 × ½ × 75/50) **[1]**.

20. (a) Mg(s) + Cu^{2+}(aq) → Mg^{2+}(aq) + Cu(s) **[1]**.
(b) Cooling line extrapolated back to 150 s gives about 31.6 °C **[1]**; ΔT = 31.6 − 19.8 = **11.8 K** (accept 11.6–12.0) **[1]**.
(c) Q = 50.0 × 4.18 × 11.8 = 2466 J = **2.47 kJ [1]**.
(d) n(Cu^{2+}) = 0.100 × 50.0/1000 = 5.00×10^{-3} mol (limiting) **[1]**; ΔH = −2.47 / 5.00×10^{-3} = **−493 kJ mol^{-1}** (accept −485 to −502 from your ΔT) **[1]**.
(e) Heat was being lost to the surroundings while the reaction was still taking place, so the highest reading is below the no-loss temperature **[1]**.

21. (a) V(H_2SO_4) = 60.0 − 40.0 = **20.0 cm^3 [1]**; at this point the reactants are in the stoichiometric ratio, so neither is in excess and the maximum amount of water forms in the fixed volume **[1]**.
(b) n(NaOH) = 1.20 × 40.0/1000 = 0.0480 mol; n(H_2SO_4) = 0.0480/2 = 0.0240 mol **[1]**; c = 0.0240 / 0.0200 = **1.20 mol dm^{-3} [1]**.
(c) Q = 60.0 × 4.18 × 10.6 = 2658 J; n(H_2O) = n(NaOH) = 0.0480 mol **[1]**; ΔH = −2.66/0.0480 = **−55.4 kJ mol^{-1} [1]**.

22. (a) Q = 150.0 × 4.18 × 18.4 = 11 537 J = **11.5 kJ [1]**.
(b) Assume the heat lost by the iron equals the heat gained by the water (no heat lost) **[1]**; 11 537 = 25.00 × 0.449 × (T − 38.4), so T − 38.4 = 1028 and T = **1.07×10^{3} °C [1]**.
(c) Higher: the bolt loses heat to the air during the transfer (and some water may boil on contact), so the heat gained by the water underestimates the heat the bolt carried **[1]**.

23. (a) Q = 50.0 × 4.18 × 6.8 = 1421 J **[1]**; n(KNO_3) = 5.05/101.11 = 0.0499 mol **[1]**; ΔH = +1.42/0.0499 = **+28.5 kJ mol^{-1}** (positive sign required) **[1]**.
(b) The solution (the products) has the greater enthalpy, because the process is endothermic **[1]**.

24. "Human error" is too vague and repeating will not fix the problem: the result is less exothermic because heat is lost to the surroundings, a **systematic** error that repeats do not remove **[1]**. Improvements, any two **[2]**: use a lid and insulate the cup (or nest two cups); record the temperature over time and extrapolate the cooling curve back to the time of mixing; use a temperature probe with data logging to capture the maximum; stir so the mixture reaches a uniform temperature quickly.

Where this fits in Reactivity 1

- **Booklet 1** (this booklet) — Reactivity 1.1.1 to 1.1.4: heat and temperature, exothermic and endothermic reactions, energy profiles, standard enthalpy changes and calorimetry.
- **Next:** Reactivity 1.2 — energy cycles: bond enthalpies, Hess's law and, for HL, formation enthalpies and Born–Haber cycles.
- Links forward: **Reactivity 1.3** (energy from fuels), **Reactivity 1.4** HL (entropy and spontaneity), **Reactivity 2.2** (activation energy and rate).

These notes are a free resource, shared for any IB student who finds them useful.

If a particular idea here still isn't clicking, specialist one-to-one support in Physics, Chemistry and Maths from a past A Level examiner is available at iBLaurel — call or message **+44 7404 710097**, or use the contact form at **www.iBLaurel.com**.

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