iBLaurel

Premium IB tuition — Physics, Chemistry, Maths

IB DIPLOMA CHEMISTRY · FIRST ASSESSMENT 2025

Reactivity 1.2
Energy Cycles in Reactions

Booklet 3 · Higher Level · Understandings 1.2.3 – 1.2.5

Standard enthalpy changes of combustion and formation, calculating ΔH from ΔH_f and ΔH_c data, and Born–Haber cycles for ionic compounds

Complete revision notes, 18 worked examples, model answers & 36 practice questions · Higher Level

How the level tags work in this booklet

HL Understandings 1.2.3, 1.2.4 and 1.2.5 are **additional higher level** material. Everything in this booklet is HL only. Standard level students need Booklets 1 and 2.

★ marks harder, multi-step questions of the kind that separate grade 6 from grade 7. This booklet builds directly on Hess's law from **Booklet 2**.

Inside this booklet

1. What the syllabus actually asks of you
2. 1.2.3 — Standard enthalpy of combustion
3. 1.2.3 — Standard enthalpy of formation
4. 1.2.4 — Calculating ΔH from ΔH_f data
5. 1.2.4 — Calculating ΔH from ΔH_c data
6. 1.2.4 — Choosing a method
7. 1.2.5 — Born–Haber cycles: the terms
8. 1.2.5 — Born–Haber calculations
9. 1.2.5 — What affects lattice enthalpy
10. Key points to memorise
11. Check yourself
12. Answers

1 What the syllabus actually asks of you

Booklet 2 introduced Hess's law with equations that were simply given to you. At higher level you need two families of standard data that make Hess's law routine. Standard enthalpies of **formation** and **combustion** are tabulated for thousands of substances, and with a single formula either set will give the enthalpy change of almost any reaction. You also need the **Born–Haber cycle**, which is Hess's law applied to the formation of an ionic solid. It is the only way to obtain experimental lattice enthalpies, and it explains why ionic compounds form even though making ions from atoms costs energy.

Understanding	What it says	What you must be able to do	Level
1.2.3 ΔH_c and ΔH_f	Standard enthalpy changes of combustion, $\Delta H_c^\ominus$, and formation, $\Delta H_f^\ominus$, data are used in thermodynamic calculations.	Deduce equations and solutions to problems involving these terms. Enthalpy of combustion and formation data are given in the data booklet.	HL
1.2.4 Calculating ΔH	An application of Hess's law uses enthalpy of formation data or enthalpy of combustion data to calculate the enthalpy change of a reaction.	Calculate enthalpy changes using $\Delta H^\ominus = \Sigma\Delta H_f^\ominus(\text{products}) - \Sigma\Delta H_f^\ominus(\text{reactants})$ and $\Delta H^\ominus = \Sigma\Delta H_c^\ominus(\text{reactants}) - \Sigma\Delta H_c^\ominus(\text{products})$. **Both equations are given in the data booklet.**	HL
1.2.5 Born–Haber	A Born–Haber cycle is an application of Hess's law, used to show energy changes in the formation of an ionic compound.	Interpret and determine values from a Born–Haber cycle for compounds composed of univalent and divalent ions. The cycle includes ionization energies, enthalpy of atomization (using sublimation and/or bond enthalpies), electron affinities, lattice enthalpy and enthalpy of formation. **The construction of a complete Born–Haber cycle will not be assessed.**	HL

Linking questions and connections for 1.2.3 – 1.2.5

- **Structure 2.2:** would you expect allotropes of an element, such as diamond and graphite, to have different $\Delta H_f^\ominus$ values? → section 3.4
- **Structure 2.1:** what are the factors that influence the strength of lattice enthalpy in an ionic compound? → section 9
- **Structure 1.3 and 3.1:** first and successive ionization energies, and electron affinities, are the steps of a Born–Haber cycle. → section 7
- **Reactivity 1.3:** comparing fuels by energy released per mole and per gram. → section 2.3
- **Reactivity 1.4:** ΔH_f data combine with entropy values to give ΔG, which decides whether a reaction is spontaneous.

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

The data booklet gives tables of $\Delta H_f^\ominus$ and $\Delta H_c^\ominus$ values, **first ionization energies** and **electron affinities**, **enthalpies of atomization**, **lattice enthalpies** (experimental and, for some compounds, theoretical) and the two ΔH equations above. It does not give the definitions, the signs of each Born–Haber step, or which multiples to use. In exams, always use the values printed in the question or your data booklet; the numbers in this booklet are typical values and may differ slightly from your copy.

2 Reactivity 1.2.3 — Standard enthalpy of combustion HL

2.1 Definition

Definition to learn

The **standard enthalpy change of combustion, $\Delta H_c^\ominus$**, of a substance is the enthalpy change when **one mole** of the substance is burned **completely** in oxygen under standard conditions (100 kPa, usually 298 K), with all reactants and products in their standard states.

"Completely" means every carbon atom ends up as $CO_2(g)$ and every hydrogen atom as $H_2O(l)$. Because water is a **liquid** at 298 K, the standard value includes the heat released when the steam condenses. Combustion is always exothermic, so $\Delta H_c^\ominus$ is always negative. Experimentally, values can be found by calorimetry (Booklet 1), and accurate values come from a bomb calorimeter.

2.2 Writing combustion equations

Method — balance in this order

1. Put **1** in front of the fuel: the definition is per mole of fuel.
2. Balance **C** with $CO_2(g)$, then **H** with $H_2O(l)$.
3. Count the O atoms on the right, subtract any O in the fuel, and balance with $O_2(g)$. Fractions such as 3½ or 9½ are normal.
4. Add state symbols for standard states.

Worked example 1 — combustion equations

Substance	Thermochemical equation for $\Delta H_c^\ominus$	$\Delta H_c^\ominus$ / kJ mol^{-1}
hexane, $C_6H_{14}(l)$	$C_6H_{14}(l) + 9\frac{1}{2}O_2(g) \rightarrow 6CO_2(g) + 7H_2O(l)$ (19 O atoms on the right)	−4163
ethanol, $C_2H_5OH(l)$	$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$ (7 O on the right, 1 in ethanol)	−1367
glucose, $C_6H_{12}O_6(s)$	$C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$	−2803
hydrogen, $H_2(g)$	$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$	−286
carbon, C(graphite)	$C(graphite) + O_2(g) \rightarrow CO_2(g)$	−394

The last two equations are worth a second look: the combustion of hydrogen is also the **formation** of water, and the combustion of carbon is also the formation of carbon dioxide. So $\Delta H_c(H_2) = \Delta H_f(H_2O, l)$ and $\Delta H_c(C) = \Delta H_f(CO_2, g)$.

2.3 Comparing fuels: per mole or per gram?

ΔH_c is quoted per mole, which makes larger molecules look better simply because they are larger. For a fuel that has to be carried, the energy released **per gram** (the specific energy) is often more useful: divide $|\Delta H_c|$ by the molar mass.

Fuel	$\Delta H_c^\ominus$ / kJ mol^{-1}	M / g mol^{-1}	energy per gram / kJ g^{-1}
hydrogen, H_2	−286	2.02	**141.6**
methane, CH_4	−891	16.05	55.5
propane, C_3H_8	−2219	44.11	50.3
octane, C_8H_{18}	−5470	114.26	47.9
ethanol, C_2H_5OH	−1367	46.08	29.7
methanol, CH_3OH	−726	32.05	22.7

Octane releases six times as much energy as methane per mole but slightly less per gram. Alcohols release less per gram than alkanes because they are already partly oxidized: the O atom adds mass without adding energy. Hydrogen is outstanding per gram but, as a gas, very poor per unit volume unless compressed or liquefied. This is developed in Reactivity 1.3.

Examiner traps in ΔH_c

- Writing the equation for 2 mol of fuel ($2C_6H_{14} + 19O_2$ …) to avoid fractions. The definition requires 1 mol of fuel.
- Showing $H_2O(g)$ instead of $H_2O(l)$, or forgetting state symbols.
- Leaving out “completely” or “in oxygen” from the definition.
- Forgetting the oxygen already inside an alcohol or sugar when balancing O_2.

3 Reactivity 1.2.3 — Standard enthalpy of formation HL

3.1 Definition and standard states

Definition to learn

The **standard enthalpy change of formation, $\Delta H_f^\ominus$**, of a substance is the enthalpy change when **one mole** of the substance is formed from its **elements in their standard states** under standard conditions (100 kPa, usually 298 K).

The **standard state** of an element is its most stable form under these conditions.

Element	Standard state at 298 K, 100 kPa	Element	Standard state at 298 K, 100 kPa
hydrogen, nitrogen, oxygen	$H_2(g)$, $N_2(g)$, $O_2(g)$ — diatomic gases	carbon	C(graphite), not diamond or C_{60}
fluorine, chlorine	$F_2(g)$, $Cl_2(g)$	sulfur, phosphorus	S(s) (or $S_8(s)$); P(s) (white, P_4)
bromine	$Br_2(\mathbf{l})$	mercury	Hg(**l**)
iodine	$I_2(\mathbf{s})$	metals (except Hg)	M(s), e.g. Na(s), Mg(s), Fe(s)

The key consequence

The standard enthalpy of formation of an **element in its standard state is zero**: forming $O_2(g)$ from $O_2(g)$ involves no change at all. Any other form of an element has a non-zero value: $\Delta H_f^\ominus$ of $O_3(g)$, $Br_2(g)$, C(diamond) and H(g) are all positive.

3.2 Writing formation equations

Write **1 mol** of the compound on the right, its elements in their standard states on the left, and balance using fractions where needed. Formation equations often describe reactions that do not actually happen; that does not matter, because the value is found indirectly using Hess's law.

Worked example 2 — formation equations

Compound	Thermochemical equation for $\Delta H_f^\ominus$	$\Delta H_f^\ominus$ / kJ mol^{-1}
ethanol, $C_2H_5OH(l)$	$2C(graphite) + 3H_2(g) + ½O_2(g) \rightarrow C_2H_5OH(l)$	−278
glucose, $C_6H_{12}O_6(s)$	$6C(graphite) + 6H_2(g) + 3O_2(g) \rightarrow C_6H_{12}O_6(s)$	−1277
calcium carbonate, $CaCO_3(s)$	$Ca(s) + C(graphite) + 1½O_2(g) \rightarrow CaCO_3(s)$	−1207
sodium chloride, NaCl(s)	$Na(s) + ½Cl_2(g) \rightarrow NaCl(s)$	−411
ammonia, $NH_3(g)$	$½N_2(g) + 1½H_2(g) \rightarrow NH_3(g)$	−46
hydrogen bromide, HBr(g)	$½H_2(g) + ½Br_2(l) \rightarrow HBr(g)$	−36
benzene, $C_6H_6(l)$	$6C(graphite) + 3H_2(g) \rightarrow C_6H_6(l)$	+49

3.3 What ΔH_f tells you about stability

Because every ΔH_f value is measured from the same zero (the elements), ΔH_f is a measure of a compound's stability **relative to its elements**. A negative value (most compounds) means the compound has lower enthalpy than its elements and is energetically more stable than them. A positive value, as for benzene (+49), ethene (+52), NO (+90) or hydrazine (+50), means the compound is energetically unstable relative to its elements, although it may still exist for a long time if the activation energy for decomposition is high.

3.4 Allotropes

Graphite and diamond have different structures (Structure 2.2), so they have different enthalpies. Graphite is the standard state, so $\Delta H_f^\ominus$(graphite) = 0 and $\Delta H_f^\ominus$(diamond) = +1.9 kJ mol^{-1}, the value found in Booklet 2 from combustion data. Graphite is more stable, partly because of the extra delocalized π-bonding between the atoms within each of its layers.

Worked example 3 — spotting zero values

Q. Which of these have a standard enthalpy change of formation of zero at 298 K and 100 kPa? H(g), Hg(l), Hg(g), C(diamond), Si(s), $I_2(s)$, $I_2(g)$, $O_3(g)$, $S_8(s)$.

A. Hg(l), Si(s), $I_2(s)$ and $S_8(s)$: each is an element in its standard state. H(g), Hg(g), $I_2(g)$, $O_3(g)$ and C(diamond) are elements but **not** in their standard states, so their values are not zero.

Model answer: "Define standard enthalpy change of formation and explain why $\Delta H_f^\ominus$ of $O_2(g)$ is zero." (3 marks)

The enthalpy change when one mole of a substance is formed from its elements **[1]** in their standard states under standard conditions of 100 kPa (and 298 K) **[1]**. $O_2(g)$ is the element in its standard state, so forming it from itself involves no chemical change and no enthalpy change **[1]**.

Examiner traps in ΔH_f

- Writing elements in the wrong state: $Br_2(g)$ instead of $Br_2(l)$, $I_2(g)$ instead of $I_2(s)$, O instead of O_2, C(diamond) instead of C(graphite).
- Forming 2 mol of product, e.g. $2Na + Cl_2 \rightarrow 2NaCl$. Formation is always per mole of the compound.
- Using a compound as a starting material: "$CO + \frac{1}{2}O_2 \rightarrow CO_2$" is not the formation of CO_2.
- Thinking a positive ΔH_f means a compound cannot exist. It means only that it is unstable relative to its elements.

4 Reactivity 1.2.4 — Calculating ΔH from ΔH_f data HL

4.1 Why the formula works

Every reactant and every product can, in principle, be made from the same set of elements. That gives a common point for a Hess cycle. Route 1 goes straight from the reactants to the products (ΔH). Route 2 goes from the reactants back to the elements (reversing their formation, $-\Sigma\Delta H_f$(reactants)) and then from the elements to the products ($+\Sigma\Delta H_f$(products)).

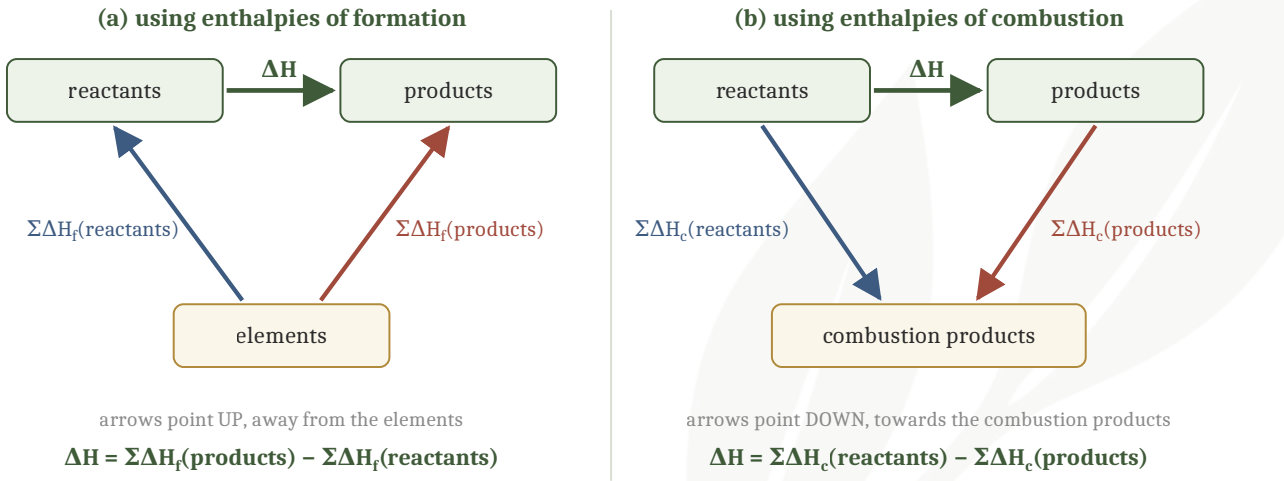


Figure 1 — The two general cycles. In (a) the arrows from the elements point up to reactants and products; in (b) the arrows point down to the combustion products. That reversal is why the two formulas are opposite ways round.

From Figure 1(a): $\Sigma\Delta H_f$(reactants) + ΔH = $\Sigma\Delta H_f$(products), so

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

"products minus reactants" · multiply each ΔH_f by its coefficient · elements in standard states count as zero

Method — four steps

1. Write the balanced equation and put each ΔH_f value under its formula.
2. Multiply each value by the coefficient in the equation. Write 0 under elements.
3. Add up the products; add up the reactants.
4. ΔH = products − reactants. Keep brackets round negative numbers.

4.2 Worked examples

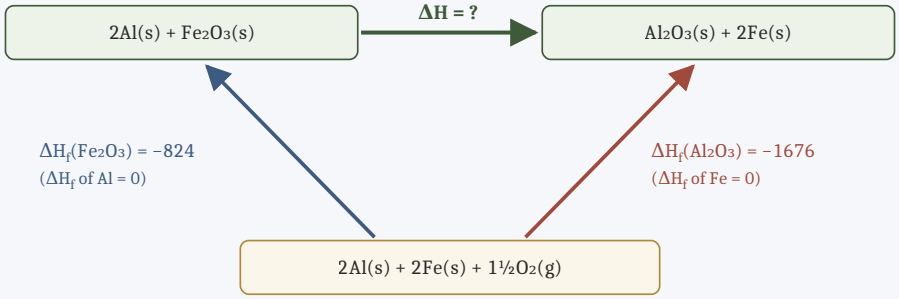
Worked example 4 — combustion of methane from formation data

	$CH_4(g)$	$+ 2O_2(g)$	$\rightarrow CO_2(g)$	$+ 2H_2O(l)$
ΔH_f / kJ mol^{-1}	−75	0	−394	−286
× coefficient	−75	0	−394	2(−286) = −572

$\Delta H = [-394 + (-572)] - [-75 + 0] = -966 + 75 =$ **−891 kJ mol^{-1}**, exactly the data booklet value of $\Delta H_c(CH_4)$.

Worked example 5 — the thermite reaction, with its cycle

Q. $2Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3(s) + 2Fe(s)$. ΔH_f: $Fe_2O_3(s)$ −824, $Al_2O_3(s)$ −1676 kJ mol^{-1}.



$\Delta H = [-1676 + 2(0)] - [2(0) + (-824)] =$ **−852 kJ mol^{-1}**

This is so exothermic that the iron is produced molten, which is why the reaction is used to weld railway lines.

Worked example 6 — oxidation of ammonia (the first step in making nitric acid)

Q. $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$. ΔH_f: $NH_3(g)$ −46, $NO(g)$ +90, $H_2O(g)$ −242 kJ mol^{-1}.

products: $4(+90) + 6(-242) = 360 - 1452 = -1092$ kJ · reactants: $4(-46) + 0 = -184$ kJ

$\Delta H = -1092 - (-184) =$ **−908 kJ** (per mole of reaction as written)

Note the state of water: the question gives ΔH_f of H_2O(**g**), so that is the value to use.

Worked example 7 — working backwards to an unknown ΔH_f

Q. $\Delta H_c^{\ominus}(C_2H_5OH, l) = -1367$ kJ mol^{-1}. Using $\Delta H_f(CO_2) = -394$ and $\Delta H_f(H_2O, l) = -286$ kJ mol^{-1}, calculate ΔH_f of ethanol.

$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$

$-1367 = [2(-394) + 3(-286)] - [\Delta H_f(C_2H_5OH) + 0] = -1646 - \Delta H_f$

$\Delta H_f(C_2H_5OH) = -1646 + 1367 =$ **−279 kJ mol^{-1}** (data booklet: −278)

Examiner traps with ΔH_f calculations

- Doing reactants − products. For formation data it is always **products − reactants**.
- Forgetting to multiply by coefficients, especially for water.
- Giving O_2, Fe, Al or other elements a non-zero value, or giving a compound a value of zero.
- Using $\Delta H_f(H_2O, l)$ when the equation shows $H_2O(g)$ (or the reverse).
- Sign slips with double negatives: write −(−184) as +184 explicitly.

5 Reactivity 1.2.4 — Calculating ΔH from ΔH_c data HL

5.1 Why the formula is the other way round

Organic compounds, carbon and hydrogen can all be burned to the same **combustion products**, $CO_2(g)$ and $H_2O(l)$. From Figure 1(b): $\Delta H + \Sigma\Delta H_c(\text{products}) = \Sigma\Delta H_c(\text{reactants})$, so

$$\Delta H^\ominus = \Sigma\Delta H_c^\ominus(\text{reactants}) - \Sigma\Delta H_c^\ominus(\text{products})$$

"reactants minus products" · O_2, CO_2 and H_2O cannot burn, so their ΔH_c counts as zero

This method is ideal for **formation reactions of organic compounds**, which cannot be carried out directly but whose combustion can be measured easily.

Worked example 8 — the enthalpy of formation of glucose

	$6C(\text{graphite})$	$+ 6H_2(g)$	$+ 3O_2(g)$	$\rightarrow C_6H_{12}O_6(s)$
ΔH_c / kJ mol^{-1}	−394	−286	0	−2803
× coefficient	−2364	−1716	0	−2803

$\Delta H_f = [-2364 + (-1716)] - [-2803] = -4080 + 2803 =$ **−1277 kJ mol^{-1}**

Worked example 9 — benzene, and a positive enthalpy of formation

Q. Calculate ΔH_f of benzene, $C_6H_6(l)$, given ΔH_c: C(graphite) −394, $H_2(g)$ −286, $C_6H_6(l)$ −3268 kJ mol^{-1}. Compare with the data booklet value of +49 kJ mol^{-1}.

$6C(\text{graphite}) + 3H_2(g) \rightarrow C_6H_6(l)$

$\Delta H_f = [6(-394) + 3(-286)] - [-3268] = -2364 - 858 + 3268 =$ **+46 kJ mol^{-1}**

The small difference arises because the answer is a small number found by subtracting two large ones (−3222 and −3268): a 0.1 % uncertainty in each ΔH_c value becomes several kJ in the answer. Benzene is unstable relative to its elements, yet it does not decompose because the activation energy is very high.

Worked example 10 — a reaction between two organic compounds

Q. Ethanol is made industrially by hydrating ethene: $C_2H_4(g) + H_2O(g) \rightarrow C_2H_5OH(l)$. Use $\Delta H_c(C_2H_4) = -1411$ and $\Delta H_c(C_2H_5OH) = -1367$ kJ mol^{-1} to estimate ΔH. Water does not burn, so its ΔH_c is zero.

$\Delta H = [-1411 + 0] - [-1367] =$ **-44 kJ mol^{-1}**

Strictly, the combustion data give the value for liquid water as a reactant; for steam the reaction is about 44 kJ more exothermic. Questions usually ask you to ignore this.

Worked example 11 — answers in algebra

Q. The ΔH_c values of ethene, hydrogen and ethane are a, b and c respectively. Which expression gives ΔH for $C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$?

A. Reactants − products = **a + b − c**. With a = −1411, b = −286 and c = −1561, $\Delta H = -136$ kJ mol^{-1}, the same value found in Booklet 2 by combining equations.

Examiner traps with ΔH_c calculations

- Using "products − reactants" out of habit. With combustion data it is **reactants − products**.
- Giving O_2, CO_2 or H_2O a ΔH_c value. They cannot burn further: zero.
- Forgetting that $\Delta H_c(C)$ and $\Delta H_c(H_2)$ must be multiplied by the number of moles of C and H_2 in the formation equation.

6 Reactivity 1.2.4 — Choosing a method HL

Data given	Formula	When to use it, and how reliable it is
ΔH_f values	$\Sigma\Delta H_f$(products) − $\Sigma\Delta H_f$(reactants)	Any reaction, including ionic solids and inorganic reactions. Accurate: values are measured for the actual substances in their actual states.
ΔH_c values	$\Sigma\Delta H_c$(reactants) − $\Sigma\Delta H_c$(products)	Reactions of substances that burn: organic compounds, C, H_2. Accurate for the same reason, though small answers from large numbers carry larger uncertainty.
bond enthalpies (Booklet 2)	ΣE(bonds broken) − ΣE(bonds formed)	Only when all species are gaseous covalent molecules. Less accurate: average values, and no intermolecular forces included.

Worked example 12 — the same reaction two ways

Q. For the hydrogenation of propene, $CH_3CH{=}CH_2(g) + H_2(g) \rightarrow CH_3CH_2CH_3(g)$, calculate ΔH (a) from ΔH_c: propene −2058, H_2 −286, propane −2219 kJ mol^{-1}; (b) from bond enthalpies: C=C 614, H–H 436, C–C 346, C–H 414 kJ mol^{-1}.

(a) $\Delta H = [-2058 + (-286)] - [-2219] =$ **−125 kJ mol^{-1}**

(b) Only the C=C and H–H bonds break; one C–C and two C–H bonds form (all other bonds are unchanged):
$\Delta H = (614 + 436) - (346 + 2 \times 414) = 1050 - 1174 =$ **−124 kJ mol^{-1}**

The agreement is unusually close because every species is a gas and the bonds involved are typical of the averages. With liquids (for example cyclohexene to cyclohexane), the bond enthalpy method would miss the energy of the intermolecular forces and the agreement would be poorer.

Model answer: "Suggest why ΔH calculated from ΔH_f values is more accurate than ΔH calculated from average bond enthalpies." (2 marks)

ΔH_f values are measured for the specific compounds in their standard states, whereas bond enthalpies are averages over many different compounds **[1]**. Bond enthalpies apply only to the gaseous state, so they ignore the enthalpy changes of intermolecular forces in liquids and solids **[1]**.

7 Reactivity 1.2.5 — Born–Haber cycles: the terms HL

7.1 Why ionic compounds form

When sodium reacts with chlorine, an electron moves from each Na atom to a Cl atom. Making the ions from gaseous atoms is actually **endothermic**:

$$Na(g) + Cl(g) \rightarrow Na^+(g) + Cl^-(g) \quad \Delta H = IE(Na) + EA(Cl) = +496 + (-349) = +147 \text{ kJ mol}^{-1}$$

Gaining a noble gas configuration is not, by itself, a reason for the reaction. The driving force is the next step: oppositely charged gaseous ions come together to form an **ionic lattice**, which releases a great deal of energy because of the strong electrostatic attraction between the ions (Structure 2.1). A **Born–Haber cycle** applies Hess's law to these steps.

7.2 The enthalpy changes in a Born–Haber cycle

Definitions to learn

Term	Definition	Example equation	Sign
Enthalpy of atomization, $\Delta H_{atom}^{\ominus}$	enthalpy change when **one mole of gaseous atoms** is formed from the element in its standard state	$Na(s) \rightarrow Na(g)$ +107 $\frac{1}{2}Cl_2(g) \rightarrow Cl(g)$ +121 $\frac{1}{2}Br_2(l) \rightarrow Br(g)$ +112	+
First ionization energy, IE_1	minimum energy needed to remove **one mole of electrons** from **one mole of gaseous atoms** to form one mole of gaseous 1+ ions	$Na(g) \rightarrow Na^+(g) + e^-$ +496	+
Second ionization energy, IE_2	energy to remove one mole of electrons from one mole of gaseous 1+ ions	$Mg^+(g) \rightarrow Mg^{2+}(g) + e^-$ +1451	+
First electron affinity, EA_1	enthalpy change when **one mole of gaseous atoms** gains one mole of electrons to form one mole of gaseous 1− ions	$Cl(g) + e^- \rightarrow Cl^-(g)$ −349	− (usually)
Second electron affinity, EA_2	enthalpy change when one mole of gaseous 1− ions gains one mole of electrons to form gaseous 2− ions	$O^-(g) + e^- \rightarrow O^{2-}(g)$ +753	+
Lattice enthalpy, $\Delta H_{lat}^{\ominus}$	enthalpy change when **one mole of a solid ionic compound** is separated into its **gaseous ions** under standard conditions	$NaCl(s) \rightarrow Na^+(g) + Cl^-(g)$ +790	+
Enthalpy of formation, $\Delta H_f^{\ominus}$	as in section 3	$Na(s) + \frac{1}{2}Cl_2(g) \rightarrow NaCl(s)$ −411	usually −

Three points that cost marks every year

- **Lattice enthalpy is endothermic in the IB.** The data booklet defines it as the lattice *breaking apart*, so values are positive. Some textbooks and A-level courses use lattice *formation* (negative). In a Born–Haber cycle the lattice-forming step is therefore $\mathbf{-\Delta H_{lat}}$.
- **Atomization of a diatomic gas is half its bond enthalpy**: $\Delta H_{atom}(Cl) = \frac{1}{2}E(Cl–Cl) = \frac{1}{2}(242) = +121$. For bromine and iodine, which are not gases in their standard states, the atomization also includes vaporization or sublimation, so $\Delta H_{atom}(Br) = +112$, not $\frac{1}{2}(193)$.
- **The second electron affinity is always endothermic.** The incoming electron is repelled by the negative charge of the 1− ion, so energy must be supplied to force it on.

Model answer: "Explain why the first electron affinity of oxygen is exothermic but the second is endothermic." (2 marks)

In the first step the electron is attracted by the nucleus of the neutral oxygen atom, so energy is released **[1]**. In the second step the electron is added to a negative ion and is repelled by its charge, so energy must be supplied to overcome the repulsion **[1]**.

8 Reactivity 1.2.5 — Born–Haber calculations HL

8.1 Sodium chloride

Lattice enthalpies cannot be measured directly. Instead, imagine forming NaCl(s) from its elements by an indirect route through gaseous ions. By Hess's law the indirect route has the same enthalpy change as the direct formation.

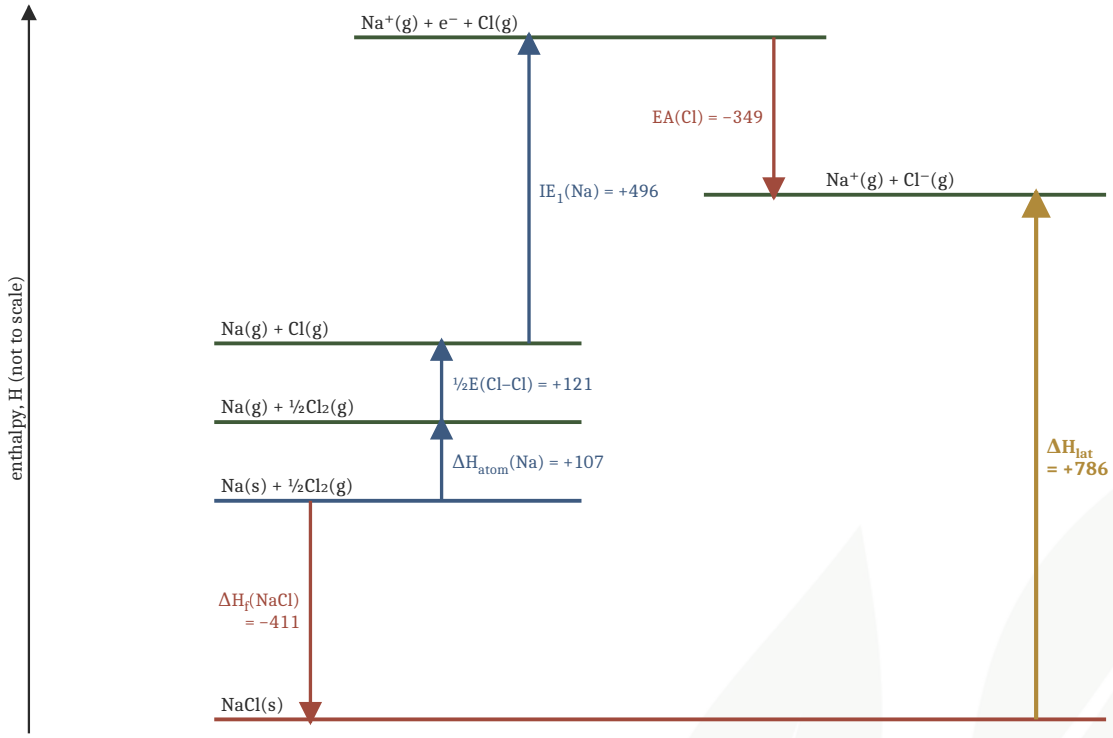


Figure 2 — Born–Haber cycle for sodium chloride. Upward arrows are endothermic, downward arrows exothermic. The gold arrow is the lattice enthalpy (lattice → gaseous ions).

Worked example 13 — the lattice enthalpy of sodium chloride

Going up from NaCl(s) to the gaseous ions directly (ΔH_{lat}) must equal going round the other way: back up to the elements ($-\Delta H_f$), then through the other steps.

$$\Delta H_{lat} = -\Delta H_f + \Delta H_{atom}(Na) + IE_1(Na) + \tfrac{1}{2}E(Cl\text{–}Cl) + EA(Cl)$$

$$\Delta H_{lat} = -(-411) + 107 + 496 + 121 + (-349) = \mathbf{+786\ kJ\ mol^{-1}}$$

The data booklet value, +790 kJ mol^{-1}, differs slightly only because of rounding in the data. Values obtained from Born–Haber cycles are called **experimental** lattice enthalpies, because every other term in the cycle is measured.

A general method that never fails

Write **ΔH_f = (sum of all the steps from the elements to the gaseous ions) − ΔH_{lat}**, put in the values, and solve for whichever term is unknown. Use the correct **multiples**:

- two atoms of the same element in the formula ($MgCl_2$, Na_2O) → double its atomization **and** its IE or EA;
- a 2+ ion → $IE_1 + IE_2$; a 2− ion → $EA_1 + EA_2$;
- $\frac{1}{2}O_2(g) \rightarrow O(g)$ is $\frac{1}{2}E(O{=}O) = +249$; $Cl_2(g) \rightarrow 2Cl(g)$ is the full $E(Cl{-}Cl) = +242$.

8.2 Magnesium oxide: a 2+/2− compound

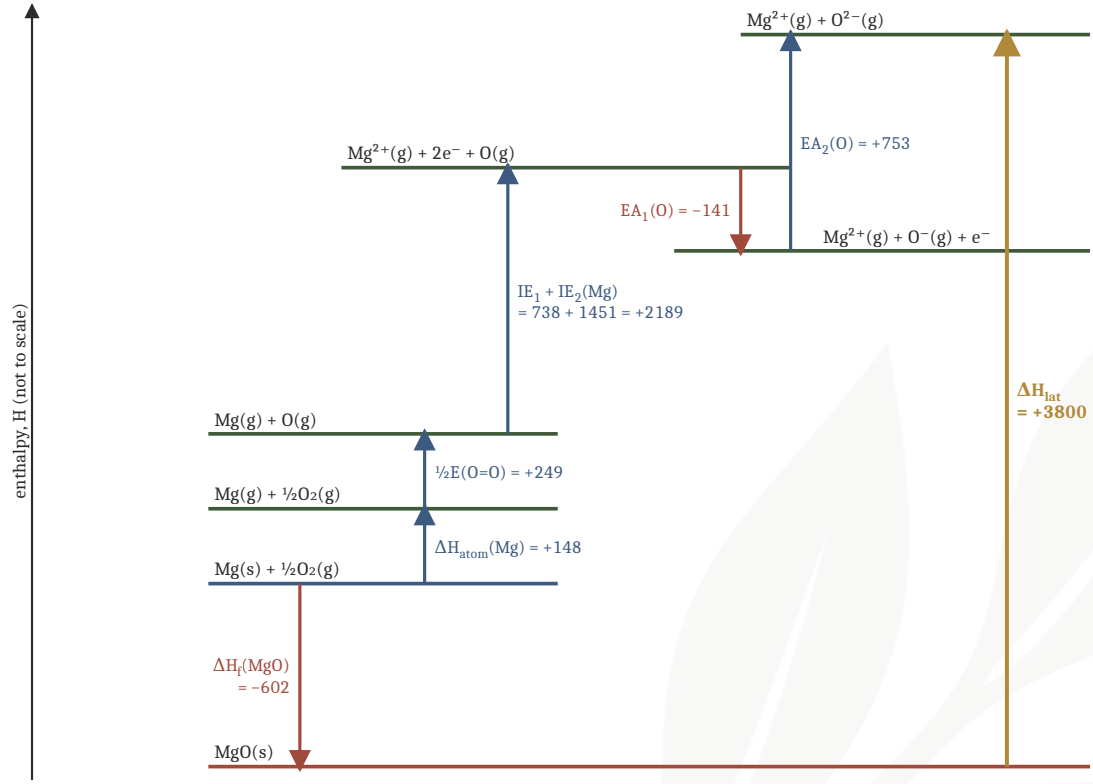


Figure 3 — Born–Haber cycle for magnesium oxide. Both ionization energies of Mg and both electron affinities of O appear; the second electron affinity points *up* because it is endothermic.

Worked example 14 — the lattice enthalpy of magnesium oxide

$$\Delta H_{lat} = -\Delta H_f + \Delta H_{atom}(Mg) + [IE_1 + IE_2](Mg) + \tfrac{1}{2}E(O{=}O) + [EA_1 + EA_2](O)$$

$$\Delta H_{lat} = 602 + 148 + (738 + 1451) + 249 + (-141 + 753) = \mathbf{+3800\ kJ\ mol^{-1}}$$

Notice how large the ionization step is (+2189 kJ) and that adding the second electron to oxygen is costly (+753 kJ). Magnesium oxide still forms, with $\Delta H_f = -602$ kJ mol^{-1}, because the lattice formed from 2+ and 2− ions releases almost five times as much energy as the NaCl lattice.

8.3 More worked examples

Worked example 15 — magnesium chloride (two chloride ions)

Q. Calculate $\Delta H_{lat}(MgCl_2)$. Data / kJ mol^{-1}: $\Delta H_f(MgCl_2)$ −642; $\Delta H_{atom}(Mg)$ +148; $IE_1(Mg)$ +738; $IE_2(Mg)$ +1451; E(Cl–Cl) +242; EA(Cl) −349.

steps from elements to gaseous ions: 148 + (738 + 1451) + 242 + 2(−349) = +1881 kJ

−642 = 1881 − ΔH_{lat} ⇒ ΔH_{lat} = 1881 + 642 = **+2523 kJ mol^{-1}**

Two moles of Cl atoms are needed: that is the **whole** bond enthalpy of Cl_2 (242, not 121), and the electron affinity must be doubled.

Worked example 16 — potassium oxide (two potassium ions)

Q. Calculate $\Delta H_{lat}(K_2O)$. Data / kJ mol^{-1}: $\Delta H_f(K_2O)$ −361; $\Delta H_{atom}(K)$ +89; $IE_1(K)$ +419; $\Delta H_{atom}(O)$ +249; $EA_1(O)$ −141; $EA_2(O)$ +753.

steps: 2(89) + 2(419) + 249 + (−141 + 753) = 178 + 838 + 249 + 612 = +1877 kJ

ΔH_{lat} = 1877 − (−361) = **+2238 kJ mol^{-1}**

Worked example 17 — finding an electron affinity

Q. The experimental lattice enthalpy of NaCl is +790 kJ mol^{-1}. Use $\Delta H_f(NaCl)$ −411, $\Delta H_{atom}(Na)$ +107, $IE_1(Na)$ +496 and $\Delta H_{atom}(Cl)$ +121 kJ mol^{-1} to calculate the electron affinity of chlorine.

−411 = 107 + 496 + 121 + EA(Cl) − 790

EA(Cl) = −411 − 107 − 496 − 121 + 790 = **−345 kJ mol^{-1}** (data booklet: −349)

Second electron affinities, such as $EA_2(O)$, cannot be measured directly. They are found in exactly this way, from Born–Haber cycles.

Worked example 18 — why is there no $NaCl_2$? ★

Q. Estimate ΔH_f of the hypothetical compound $NaCl_2$, containing Na^{2+} ions, given $IE_2(Na)$ = +4562 kJ mol^{-1} and an estimated lattice enthalpy of +2500 kJ mol^{-1}. Use the other data from worked example 13.

$\Delta H_f = \Delta H_{atom}(Na) + IE_1 + IE_2 + E(Cl–Cl) + 2EA(Cl) - \Delta H_{lat}$

ΔH_f = 107 + 496 + 4562 + 242 + 2(−349) − 2500 = **+2209 kJ mol^{-1}**

Removing a second electron from sodium means breaking into the full 2p subshell, close to the nucleus. Even the larger lattice enthalpy of a 2+ ion cannot repay this, so $NaCl_2$ would be hugely unstable relative to its elements and does not form.

Examiner traps in Born–Haber cycles

- Using the sign of lattice enthalpy from a non-IB source. In IB questions ΔH_{lat} is **positive**; the lattice-forming arrow is $-\Delta H_{lat}$.
- Using the full bond enthalpy for $\frac{1}{2}Cl_2$, or half of it for $Cl_2 \rightarrow 2Cl$.
- Forgetting to double the atomization and EA (for $MgCl_2$) or the atomization and IE (for Na_2O).
- Using only IE_1 for a 2+ ion, or only EA_1 for O^{2-}.
- Writing ionization as $Na(s) \rightarrow Na^+(g)$: ionization energies start from **gaseous** atoms, so atomization must come first.

9 Reactivity 1.2.5 — What affects lattice enthalpy HL

9.1 Charge and size

The lattice enthalpy measures the strength of the electrostatic attraction between the ions. From Coulomb's law, the attraction is stronger when the charges are larger and when the ions are closer together. Lattice enthalpy is therefore:

- **proportional to the product of the ionic charges**, so MgO (2+ × 2−) is far larger than NaCl (1+ × 1−); and
- **inversely proportional to the sum of the ionic radii**, so it falls as the ions get larger down a group.

Small, highly charged ions (high **charge density**) give the largest lattice enthalpies.

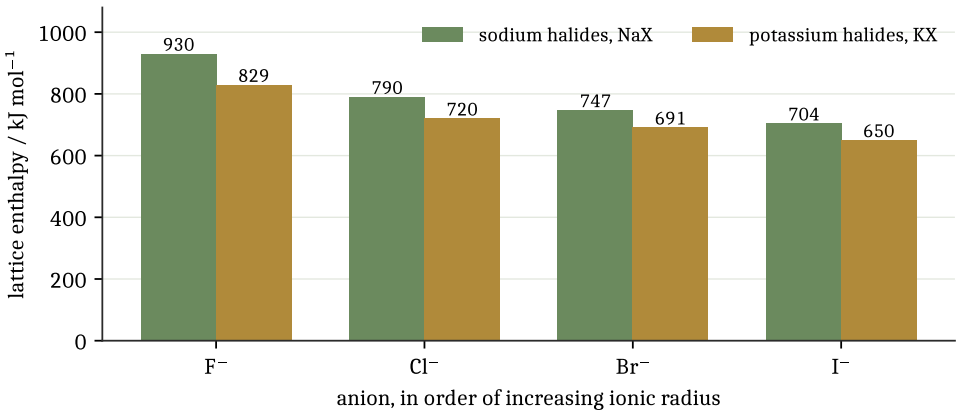


Figure 4 — Experimental lattice enthalpies of sodium and potassium halides. Each value falls as the anion gets larger, and every potassium salt is below the corresponding sodium salt because K^+ is larger than Na^+.

Compound	ions	product of charges	$r_+ + r_-$ / pm	ΔH_{lat} (experimental) / kJ mol^{-1}
NaCl	$Na^+ Cl^-$	1	102 + 181 = 283	790
KCl	$K^+ Cl^-$	1	138 + 181 = 319	720
NaF	$Na^+ F^-$	1	102 + 133 = 235	930
CaO	$Ca^{2+} O^{2-}$	4	100 + 140 = 240	3401
MgO	$Mg^{2+} O^{2-}$	4	72 + 140 = 212	3791

Ionic radii are typical values for 6-coordination; lattice enthalpies are typical experimental values. Use your data booklet values in exams.

9.2 Experimental and theoretical values: covalent character

A **theoretical** lattice enthalpy can be calculated from the ionic model, by adding up all the electrostatic attractions and repulsions in a perfect lattice of spherical point-charge ions. An **experimental** lattice enthalpy comes from a Born–Haber cycle. When the two agree closely, the bonding is almost purely ionic. When the experimental value is significantly **larger**, the bonding has some **covalent character**: a small or highly charged cation distorts (polarizes) the electron cloud of a large anion, so electron density is partly shared and the bonding is stronger than the ionic model predicts.

Compound	experimental / kJ mol^{-1}	theoretical / kJ mol^{-1}	difference	Interpretation
NaCl	790	770	2.6 %	almost purely ionic
AgCl	915	833	9.8 %	significant covalent character
AgI	889	778	14.3 %	more covalent: the large I^- ion is easily polarized

Percentage difference = (experimental − theoretical) / theoretical × 100.

Model answer: "Explain why the lattice enthalpy of magnesium oxide is much larger than that of sodium chloride." (2 marks)

Mg^{2+} and O^{2-} have twice the charges of Na^+ and Cl^-, so the product of the charges is four times as large **[1]**. The ions in MgO are also smaller, so they are closer together and the electrostatic attraction between them is stronger **[1]**.

Examiner traps on lattice enthalpy

- Explaining lattice enthalpy with "stronger bonds" or electronegativity. Use **ionic charge**, **ionic radius** and **electrostatic attraction**.
- Comparing atomic radii instead of **ionic** radii.
- Saying that a larger experimental value means the compound is "more ionic". It means the opposite: the extra strength comes from covalent character.

10 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

- **$\Delta H_c^\ominus$:** enthalpy change when one mole of a substance is completely burned in oxygen under standard conditions.
- **$\Delta H_f^\ominus$:** enthalpy change when one mole of a substance is formed from its elements in their standard states under standard conditions. Elements in their standard states: zero.
- **Standard state:** the most stable form of an element at 100 kPa and the stated temperature (usually 298 K).
- **ΔH_{atom}:** one mole of gaseous atoms formed from the element in its standard state.
- **IE_1:** one mole of electrons removed from one mole of gaseous atoms. **EA_1:** one mole of gaseous atoms gains one mole of electrons.
- **ΔH_{lat}:** one mole of a solid ionic compound separated into its gaseous ions (endothermic in the IB).

B Formulas and rules

- $\Delta H = \Sigma\Delta H_f(\text{products}) - \Sigma\Delta H_f(\text{reactants})$ · $\Delta H = \Sigma\Delta H_c(\text{reactants}) - \Sigma\Delta H_c(\text{products})$. Multiply by coefficients.
- $\Delta H_c(H_2) = \Delta H_f(H_2O, l)$; $\Delta H_c(C, \text{graphite}) = \Delta H_f(CO_2, g)$.
- Born–Haber: $\Delta H_f = \Delta H_{atom}(\text{metal}) + \text{IEs} + \Delta H_{atom}(\text{non-metal}) + \text{EAs} - \Delta H_{lat}$.
- Diatomic gas: $\Delta H_{atom} = \frac{1}{2}E(X\text{–}X)$. Double atomization and IE/EA when the formula has two of that ion.
- Atomization, IE, EA_2 and ΔH_{lat} are endothermic; EA_1 is (usually) exothermic.

C Explanations you will be asked for

- Negative ΔH_f: stable relative to the elements. Graphite (0) is more stable than diamond (+1.9).
- $\Delta H_f/\Delta H_c$ methods are more accurate than bond enthalpies: specific measured values, correct states.
- EA_2 is endothermic because the electron is repelled by the negative ion.
- Ionic compounds form because the lattice enthalpy outweighs the energy needed to make the gaseous ions.
- Lattice enthalpy increases with ionic charge and decreases with ionic radius. Experimental > theoretical indicates covalent character.

11 Check yourself

Use the data given in each question; values are in kJ mol^{-1} unless stated. Marks are shown in brackets; allow roughly 1.2 minutes per mark. Full answers with mark-scheme wording are in section 12.

Part A — Multiple choice HL

1. Which equation represents the standard enthalpy change of combustion of ethanol?

A $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$
B $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$
C $2C_2H_5OH(l) + 6O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$
D $C_2H_5OH(l) + 2O_2(g) \rightarrow 2CO(g) + 3H_2O(l)$

2. Which has a standard enthalpy change of formation of zero at 298 K and 100 kPa?

A $Br_2(g)$ **B** $Br_2(l)$ **C** $O_3(g)$ **D** C(diamond)

3. Which equation represents the standard enthalpy change of formation of calcium carbonate?

A $CaO(s) + CO_2(g) \rightarrow CaCO_3(s)$
B $Ca(s) + C(diamond) + 1½O_2(g) \rightarrow CaCO_3(s)$
C $Ca(s) + C(graphite) + 1½O_2(g) \rightarrow CaCO_3(s)$
D $Ca(s) + C(graphite) + 3O(g) \rightarrow CaCO_3(s)$

4. Which equation represents both the standard enthalpy of combustion of one substance and the standard enthalpy of formation of another?

A $H_2(g) + ½O_2(g) \rightarrow H_2O(l)$
B $CO(g) + ½O_2(g) \rightarrow CO_2(g)$
C $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$
D $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

5. ΔH_f: $SO_2(g)$ −297, $SO_3(g)$ −396. What is ΔH for $SO_2(g) + ½O_2(g) \rightarrow SO_3(g)$?

A −99 **B** +99 **C** −693 **D** −396

6. ΔH_f: $H_2O_2(l)$ −188, $H_2O(l)$ −286. What is ΔH for $2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$?

A +98 **B** −98 **C** +196 **D** −196

7. The ΔH_f values of $C_3H_8(g)$, $CO_2(g)$ and $H_2O(l)$ are x, y and z. Which expression gives ΔH_c of propane?

A x − 3y − 4z **B** 3y + 4z − x **C** x + 3y + 4z **D** 3y + 4z + x

8. Cracking: $C_{10}H_{22}(l) \rightarrow C_8H_{18}(l) + C_2H_4(g)$. ΔH_c: $C_{10}H_{22}$ −6778, C_8H_{18} −5470, C_2H_4 −1411. What is ΔH?

A −10 837 **B** −103 **C** +103 **D** +13 659

9. Which fuel releases the most energy per gram? (ΔH_c / kJ mol^{-1} and M / g mol^{-1} given.)

A H_2: −286, 2.02 **B** CH_4: −891, 16.05 **C** C_8H_{18}: −5470, 114.26 **D** C_2H_5OH: −1367, 46.08

10. ΔH_c: C(graphite) −393.5, C(diamond) −395.4. What is $\Delta H_f^{\ominus}$ of diamond?

A 0 **B** −1.9 **C** +1.9 **D** −788.9

11. Which equation represents the first electron affinity of chlorine?

A $\frac{1}{2}Cl_2(g) + e^- \rightarrow Cl^-(g)$
B $Cl(g) + e^- \rightarrow Cl^-(g)$
C $Cl(g) \rightarrow Cl^+(g) + e^-$
D $Cl_2(g) + 2e^- \rightarrow 2Cl^-(g)$

12. Which step in the Born–Haber cycle for sodium chloride is exothermic?

A $Na(s) \rightarrow Na(g)$
B $Na(g) \rightarrow Na^+(g) + e^-$
C $Cl(g) + e^- \rightarrow Cl^-(g)$
D $\frac{1}{2}Cl_2(g) \rightarrow Cl(g)$

13. Which equation represents the lattice enthalpy of magnesium chloride (IB convention)?

A $Mg^{2+}(g) + 2Cl^-(g) \rightarrow MgCl_2(s)$
B $MgCl_2(s) \rightarrow Mg^{2+}(g) + 2Cl^-(g)$
C $MgCl_2(s) \rightarrow Mg(g) + Cl_2(g)$
D $MgCl_2(s) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq)$

14. Which equation represents the standard enthalpy change of atomization of bromine?

A $Br_2(l) \rightarrow 2Br(g)$
B $Br_2(g) \rightarrow 2Br(g)$
C $\frac{1}{2}Br_2(g) \rightarrow Br(g)$
D $\frac{1}{2}Br_2(l) \rightarrow Br(g)$

15. Which compound has the largest lattice enthalpy?

A KCl
B NaF
C CaO
D MgO

16. Which expression gives the lattice enthalpy of NaCl?

A $\Delta H_f + \Delta H_{atom}(Na) + IE(Na) + \Delta H_{atom}(Cl) + EA(Cl)$
B $-\Delta H_f + \Delta H_{atom}(Na) + IE(Na) + \Delta H_{atom}(Cl) + EA(Cl)$
C $-\Delta H_f - \Delta H_{atom}(Na) - IE(Na) - \Delta H_{atom}(Cl) - EA(Cl)$
D $\Delta H_f - \Delta H_{atom}(Na) - IE(Na) - \Delta H_{atom}(Cl) + EA(Cl)$

17. For KCl: $\Delta H_{atom}(K)$ +89, IE(K) +419, $\Delta H_{atom}(Cl)$ +121, EA(Cl) −349, ΔH_{lat} +718. What is $\Delta H_f(KCl)$?

A −438
B +438
C −998
D +280

18. Why is the second electron affinity of oxygen endothermic?

A The oxygen atom is very small.
B An electron is added to a negative ion, which repels it.
C Oxygen has a high electronegativity.
D The electron enters a new shell further from the nucleus.

19. Which compound shows the greatest percentage difference between its experimental and theoretical lattice enthalpies?

A NaF
B NaCl
C KCl
D AgI

20. In the Born–Haber cycle for Na_2O, which term correctly represents the sodium steps?

A $\Delta H_{atom}(Na) + IE_1(Na)$
B $2\Delta H_{atom}(Na) + IE_1(Na)$
C $2[\Delta H_{atom}(Na) + IE_1(Na)]$
D $\Delta H_{atom}(Na) + IE_1(Na) + IE_2(Na)$

Part B — Structured questions HL

21. Write thermochemical equations, with state symbols, for: [6]

(a) the standard enthalpy of combustion of propan-1-ol, $C_3H_7OH(l)$;

(b) the standard enthalpy of combustion of octane, $C_8H_{18}(l)$;

(c) the standard enthalpy of formation of ethanoic acid, $CH_3COOH(l)$;

(d) the standard enthalpy of formation of ammonium nitrate, $NH_4NO_3(s)$;

(e) the standard enthalpy of formation of iodine monochloride, $ICl(s)$;

(f) the lattice enthalpy of calcium fluoride.

22. Use these ΔH_f values (kJ mol^{-1}) to answer the questions: $CaCO_3(s)$ −1207, $CaO(s)$ −635, $CO_2(g)$ −394, $CO(g)$ −111, $Fe_2O_3(s)$ −824, $CH_3OH(l)$ −239, $H_2O(l)$ −286, $MgO(s)$ −602.

(a) Calculate ΔH for the thermal decomposition $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$. [2]

(b) Calculate ΔH for the blast furnace reaction $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$. [2]

(c) Calculate ΔH_c of methanol and compare it with the data booklet value of −726 kJ mol^{-1}. [2]

(d) Calculate ΔH for $2MgO(s) + C(\text{graphite}) \rightarrow 2Mg(s) + CO_2(g)$ and comment on using carbon to extract magnesium. [2]

23. Hydrazine burns: $N_2H_4(l) + O_2(g) \rightarrow N_2(g) + 2H_2O(l)$, $\Delta H = -622$ kJ mol^{-1}.

(a) Calculate ΔH_f of hydrazine, using $\Delta H_f(H_2O, l) = -286$ kJ mol^{-1}. [2]

(b) Comment on the stability of hydrazine relative to its elements. [1]

24. Sodium peroxide is used in breathing apparatus: $2Na_2O_2(s) + 2CO_2(g) \rightarrow 2Na_2CO_3(s) + O_2(g)$.

(a) Calculate ΔH using ΔH_f: $Na_2O_2(s)$ −510.9, $Na_2CO_3(s)$ −1130.7, $CO_2(g)$ −394 kJ mol^{-1}. [3]

(b) Outline why bond enthalpies could not be used for this calculation. [1]

25. ΔH_c / kJ mol^{-1}: $C(\text{graphite})$ −394, $H_2(g)$ −286, $C_6H_{14}(l)$ −4163, $C_4H_{10}(g)$ −2878, $C_6H_{12}O_6(aq)$ −2803, $C_2H_5OH(l)$ −1367.

(a) Calculate ΔH_f of hexane. [2]

(b) Calculate ΔH_f of butane. [2]

(c) Calculate ΔH for the fermentation $C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(l) + 2CO_2(g)$. [2]

(d) Explain why ΔH_f of hexane cannot be measured directly. [1]

26. ★ Two students calculated ΔH for $C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$. Sam used average bond enthalpies (C=C 614, H–H 436, C–C 346, C–H 414); Priya used ΔH_c values (C_2H_4 −1411, H_2 −286, C_2H_6 −1561).

(a) Calculate Sam's value. [2]

(b) Calculate Priya's value. [1]

(c) State, with a reason, which value is more reliable. [1]

(d) Predict, with a reason, whether the two methods would agree more or less closely for $C_6H_{10}(l) + H_2(g) \rightarrow C_6H_{12}(l)$, and state what extra data the bond enthalpy method would need. [2]

27. Use the data in section 2.3.

(a) Show that methane releases more energy per gram than octane, although octane releases more per mole. [2]

(b) Explain why ethanol releases less energy per gram than the alkanes. [1]

(c) Hydrogen has the highest energy per gram. Suggest one practical disadvantage of hydrogen as a vehicle fuel. [1]

28. ★ $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(l)$. Using ΔH_f: $NH_3(g)$ −46, $NO(g)$ +90, $H_2O(l)$ −286 kJ mol^{-1}:

(a) calculate ΔH; [2]

(b) explain the difference from worked example 6 (−908 kJ) and calculate the enthalpy of vaporization of water from the two answers. [2]

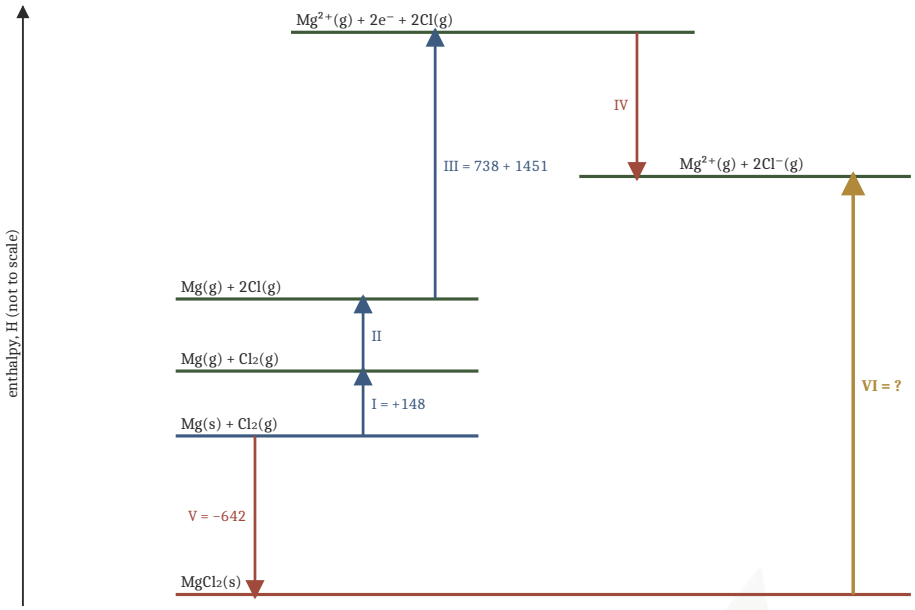
29. Potassium chloride. Data / kJ mol^{-1}: ΔH_f(KCl) −437, ΔH_{atom}(K) +89, IE_1(K) +419, E(Cl–Cl) +242, EA(Cl) −349.

(a) Write equations, with state symbols, for the atomization of chlorine and the first ionization energy of potassium. [2]

(b) Calculate the lattice enthalpy of KCl. [2]

(c) Explain why the lattice enthalpy of KCl is smaller than that of NaCl. [1]

30. The Born–Haber cycle for magnesium chloride is shown (values in kJ mol^{-1}).



(a) Identify the enthalpy changes labelled I, V and VI. [2]

(b) Calculate the values of II and IV, given E(Cl–Cl) = 242 and EA(Cl) = −349 kJ mol^{-1}. [2]

(c) Calculate the lattice enthalpy, VI. [2]

(d) The theoretical lattice enthalpy of $MgCl_2$ is +2326 kJ mol^{-1}. Explain the difference. [2]

31. Calcium fluoride, CaF_2. Data / kJ mol^{-1}: ΔH_f −1220, ΔH_{atom}(Ca) +178, IE_1(Ca) +590, IE_2(Ca) +1145, E(F–F) +159, EA(F) −328. Calculate the lattice enthalpy of CaF_2. [3]

32. Sodium oxide, Na_2O. Data / kJ mol^{-1}: ΔH_{atom}(Na) +107, IE_1(Na) +496, ΔH_{atom}(O) +249, EA_1(O) −141, EA_2(O) +753, ΔH_f(Na_2O) −414.

(a) Calculate the enthalpy changes for $Na(s) \rightarrow Na^+(g) + e^-$ and $\frac{1}{2}O_2(g) + 2e^- \rightarrow O^{2-}(g)$. [2]

(b) Calculate the lattice enthalpy of Na_2O. [2]

(c) Explain why K_2O (+2238) has a smaller lattice enthalpy than Na_2O. [1]

33. ★ The second electron affinity of oxygen cannot be measured directly. For calcium oxide: ΔH_f −635, ΔH_{atom}(Ca) +178, IE_1 + IE_2(Ca) +1735, ΔH_{atom}(O) +249, EA_1(O) −141, ΔH_{lat}(CaO) +3401 kJ mol^{-1}. Calculate EA_2(O) and explain its sign. [3]

34. Lattice enthalpies (kJ mol^{-1}): NaF 930, NaCl 790, NaBr 747, NaI 704, MgO 3791, CaO 3401.

(a) Explain the trend from NaF to NaI. [2]

(b) Explain why MgO has a larger lattice enthalpy than CaO. [1]

(c) Explain why MgO has a much larger lattice enthalpy than NaF, although the ions have similar sizes. [1]

35. ★ Experimental and theoretical lattice enthalpies (kJ mol^{-1}): NaCl 790 and 770; AgCl 915 and 833; AgI 889 and 778.

(a) Calculate the percentage difference for each compound. [2]

(b) State the assumptions of the ionic model used to calculate theoretical values. [1]

(c) Explain why the difference is largest for silver iodide. [2]

36. ★ Magnesium could, in principle, form MgCl (containing Mg^{+}) or $MgCl_2$. The estimated lattice enthalpy of MgCl is +760 kJ mol^{-1}.

(a) Use the data from question 30 to calculate ΔH_f of MgCl. **[2]**

(b) Use the values of ΔH_f to explain why magnesium chloride is $MgCl_2$ rather than MgCl, even though forming Mg^{2+} requires an extra 1451 kJ mol^{-1}. **[2]**

Part A: 20 marks · Part B: 79 marks · Total: 99 marks (about 2 hours)

12 Answers

Part A

1. A. One mole of fuel, complete combustion, liquid water. C is for 2 mol; D is incomplete.

2. B. Bromine is a liquid at 298 K, so $Br_2(l)$ is its standard state.

3. C. Elements in standard states: carbon as graphite, oxygen as $O_2(g)$.

4. A. Combustion of 1 mol H_2 and formation of 1 mol $H_2O(l)$. B is not a formation (CO is a compound); C forms 2 mol.

5. A. $-396 - (-297 + 0) = -99$.

6. D. $2(-286) - 2(-188) = -572 + 376 = -196$.

7. B. Products − reactants = $3y + 4z - x$.

8. C. Reactants − products = $-6778 - (-5470 - 1411) = +103$.

9. A. $286/2.02 = 142$ kJ g^{-1}, compared with 55.5, 47.9 and 29.7.

10. C. C(graphite) → C(diamond): reactants − products = $-393.5 - (-395.4) = +1.9$.

11. B. One mole of gaseous atoms gaining one mole of electrons.

12. C. The first electron affinity of chlorine is −349. A, B and D are endothermic.

13. B. Solid separated into gaseous ions (endothermic). A is the reverse; D is dissolving.

14. D. One mole of gaseous atoms from the element in its standard state (liquid bromine).

15. D. 2+/2− ions, and Mg^{2+} is smaller than Ca^{2+}.

16. B. $\Delta H_{lat} = -\Delta H_f +$ (all steps from the elements to the gaseous ions).

17. A. $\Delta H_f = 89 + 419 + 121 - 349 - 718 = -438$.

18. B.

19. D. The large, polarizable I^- ion with Ag^+ gives significant covalent character; the others are close to purely ionic.

20. C. Two Na atoms must each be atomized and ionized.

Part B

21. One mark each (correct species, balancing for 1 mol, state symbols):

(a) $C_3H_7OH(l) + 4½O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$

(b) $C_8H_{18}(l) + 12½O_2(g) \rightarrow 8CO_2(g) + 9H_2O(l)$

(c) $2C(graphite) + 2H_2(g) + O_2(g) \rightarrow CH_3COOH(l)$

(d) $N_2(g) + 2H_2(g) + 1½O_2(g) \rightarrow NH_4NO_3(s)$

(e) $½I_2(s) + ½Cl_2(g) \rightarrow ICl(s)$

(f) $CaF_2(s) \rightarrow Ca^{2+}(g) + 2F^-(g)$

22. (a) (−635 − 394) − (−1207) **[1]** = **+178 kJ mol^{-1} [1]**.
(b) [2(0) + 3(−394)] − [−824 + 3(−111)] = −1182 + 1157 **[1]** = **−25 kJ mol^{-1} [1]**.
(c) $CH_3OH(l) + 1½O_2 \rightarrow CO_2 + 2H_2O(l)$: [−394 + 2(−286)] − (−239) = **−727 kJ mol^{-1} [1]**; agrees with −726 to within rounding **[1]**.
(d) −394 − 2(−602) = **+810 kJ [1]**; strongly endothermic, so carbon is not a suitable reducing agent for magnesium oxide (magnesium is extracted by electrolysis instead) **[1]**.

23. (a) −622 = [0 + 2(−286)] − [ΔH_f + 0] **[1]**; ΔH_f = −572 + 622 = **+50 kJ mol^{-1} [1]**.
(b) Positive, so hydrazine is energetically unstable relative to N_2 and H_2 (it has higher enthalpy than its elements) **[1]**.

24. (a) products: 2(−1130.7) + 0 = −2261.4; reactants: 2(−510.9) + 2(−394) = −1809.8 **[1]**; ΔH = −2261.4 + 1809.8 **[1]** = **−451.6 kJ [1]**.
(b) Na_2O_2 and Na_2CO_3 are ionic solids: bond enthalpies apply only to covalent bonds in gaseous molecules **[1]**.

25. (a) $6C + 7H_2 \rightarrow C_6H_{14}$: [6(−394) + 7(−286)] − (−4163) **[1]** = −2364 − 2002 + 4163 = **−203 kJ mol^{-1} [1]**.
(b) [4(−394) + 5(−286)] − (−2878) **[1]** = −1576 − 1430 + 2878 = **−128 kJ mol^{-1} [1]**.
(c) −2803 − [2(−1367) + 0] **[1]** = **−69 kJ mol^{-1} [1]**.
(d) Carbon and hydrogen do not react directly to give hexane (a mixture of many products would form, if any reaction occurred) **[1]**.

26. (a) bonds broken C=C + H–H = 1050; bonds formed C–C + 2C–H = 1174 **[1]**; ΔH = **−124 kJ mol^{-1} [1]**.
(b) −1411 − 286 + 1561 = **−136 kJ mol^{-1} [1]**.
(c) Priya's: ΔH_c values are measured for the specific substances, whereas bond enthalpies are averages **[1]**.
(d) Less closely: cyclohexene and cyclohexane are liquids, so the bond enthalpy method ignores intermolecular forces **[1]**; it would need the enthalpies of vaporization of cyclohexene and cyclohexane **[1]**.

27. (a) methane: 891/16.05 = 55.5 kJ g^{-1}; octane: 5470/114.26 = 47.9 kJ g^{-1} **[1]**; octane has the larger $|\Delta H_c|$ per mole (5470 > 891) but the smaller value per gram **[1]**.
(b) Ethanol is already partly oxidized; its O atom adds mass but releases no energy on combustion **[1]**.
(c) Hydrogen is a gas with a very low density, so it needs heavy high-pressure or low-temperature tanks (or: it is flammable/explosive; most hydrogen is made from fossil fuels) **[1]**.

28. (a) [4(90) + 6(−286)] − [4(−46)] = 360 − 1716 + 184 **[1]** = **−1172 kJ [1]**.
(b) Forming liquid water releases extra energy as the 6 mol of steam condense **[1]**; (−1172 − (−908))/6 = −44 kJ mol^{-1} released on condensing, so $\Delta H_{vap}(H_2O)$ = **+44 kJ mol^{-1} [1]**.

29. (a) $½Cl_2(g) \rightarrow Cl(g)$ **[1]**; $K(g) \rightarrow K^+(g) + e^-$ **[1]**.
(b) −437 = 89 + 419 + 121 − 349 − ΔH_{lat} **[1]**; ΔH_{lat} = **+717 kJ mol^{-1} [1]**.
(c) K^+ is larger than Na^+, so the ions are further apart and the electrostatic attraction is weaker **[1]**.

30. (a) I: enthalpy of atomization of magnesium; V: enthalpy of formation of $MgCl_2$; VI: lattice enthalpy of $MgCl_2$ (all three for **[2]**; two for **[1]**).
(b) II = E(Cl–Cl) = **+242** (two Cl atoms) **[1]**; IV = 2 × EA(Cl) = **−698 [1]**.
(c) VI = 642 + 148 + 242 + 2189 − 698 **[1]** = **+2523 kJ mol^{-1} [1]**.
(d) The experimental value is larger **[1]** because the bonding has some covalent character: the small, highly charged Mg^{2+} ion polarizes the electron clouds of the chloride ions **[1]**.

31. Steps: 178 + 590 + 1145 + 159 + 2(−328) = +1416 kJ **[1]** (E(F–F) in full and EA doubled, for two F atoms); −1220 = 1416 − ΔH_{lat} **[1]**; ΔH_{lat} = **+2636 kJ mol^{-1} [1]**.

32. (a) $Na(s) \rightarrow Na^+(g)$: 107 + 496 = **+603 kJ mol^{-1} [1]**; $\frac{1}{2}O_2 \rightarrow O^{2-}(g)$: 249 − 141 + 753 = **+861 kJ mol^{-1} [1]**.
(b) −414 = 2(603) + 861 − ΔH_{lat} **[1]**; ΔH_{lat} = 1206 + 861 + 414 = **+2481 kJ mol^{-1} [1]**.
(c) K^+ has a larger ionic radius than Na^+ (same charge), so the attraction to O^{2-} is weaker **[1]**.

33. −635 = 178 + 1735 + 249 − 141 + EA_2 − 3401 **[1]**; EA_2 = **+745 kJ mol^{-1} [1]**; positive (endothermic) because the electron is added to a negative O^- ion and must overcome repulsion **[1]**.

34. (a) The halide ion gets larger from F^- to I^- **[1]**, so the distance between the ion centres increases and the electrostatic attraction weakens **[1]**.
(b) Mg^{2+} is smaller than Ca^{2+} (same charge), so it is closer to the O^{2-} ions **[1]**.
(c) The ionic charges are 2+ and 2− rather than 1+ and 1−, so the product of the charges is four times greater **[1]**.

35. (a) NaCl (790 − 770)/770 = **2.6 %**; AgCl 82/833 = **9.8 %**; AgI 111/778 = **14.3 %** (all three **[2]**; two **[1]**).
(b) The ions are perfect spheres (point charges) with no distortion of their electron clouds, and the bonding is 100 % ionic **[1]**.
(c) The I^- ion is large, so its outer electrons are easily distorted (polarized) by the Ag^+ ion **[1]**; electron density is partly shared between the ions, giving covalent character, which makes the lattice stronger than the ionic model predicts **[1]**.

36. (a) $\Delta H_f(MgCl)$ = 148 + 738 + 121 + (−349) − 760 **[1]** = **−102 kJ mol^{-1} [1]**.
(b) $\Delta H_f(MgCl_2)$ = −642 kJ mol^{-1}, far more exothermic than MgCl, so $MgCl_2$ is the more stable product **[1]**; the much larger lattice enthalpy of a lattice with Mg^{2+} ions (+2523 against +760) more than repays the extra ionization energy **[1]**.

Where this fits in Reactivity 1

- **Booklet 1** — Reactivity 1.1: heat, exothermic and endothermic reactions, energy profiles and calorimetry.
- **Booklet 2** — Reactivity 1.2.1 to 1.2.2: bond enthalpies and Hess's law.
- **Booklet 3** (this booklet) — HL Reactivity 1.2.3 to 1.2.5: ΔH_f, ΔH_c and Born–Haber cycles.
- **Next:** Reactivity 1.3 — energy from fuels; then HL Reactivity 1.4 — entropy and spontaneity.

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**.

Written and produced independently by iBLaurel. Not affiliated with, endorsed by, or published in association with the International Baccalaureate Organization. Syllabus references are to the IB Chemistry guide, first assessment 2025. Numerical data are typical literature values consistent with the IB Chemistry data booklet; always use the values given in examination questions.