

Reactivity 1.2 · Energy cycles in reactions

International Baccalaureate · IB Diploma · Chemistry · SL

Scope and prerequisites

This handout teaches Reactivity 1.2, Energy cycles in reactions, for the course first assessed in 2025.

Prerequisites: enthalpy change, exothermic and endothermic reactions, and calorimetry from Reactivity 1.1; covalent bonds from Structure 2.2; lattice enthalpy and ionization energy from Structures 2.1 and 1.3. You will use bond enthalpies and enthalpy data from the data booklet.

Bond enthalpy

■ Where does the energy of a reaction come from?

Burning one mole of methane in a gas cooker releases about 800 kJ. No atoms are created or destroyed, so the energy must come from the bonds: the bonds in the products are stronger than the bonds in the reactants. This unit shows how to calculate enthalpy changes from bond data and from **energy cycles** when they cannot be measured directly.

■ Breaking and making bonds

- Breaking a bond **needs** energy: bond breaking is **endothermic**.
- Making a bond **releases** energy: bond making is **exothermic**.

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

The value for a bond such as C–H differs a little from one molecule to another, so the data booklet gives an **average bond enthalpy**: the mean over many compounds.

■ Calculating ΔH from bond enthalpies

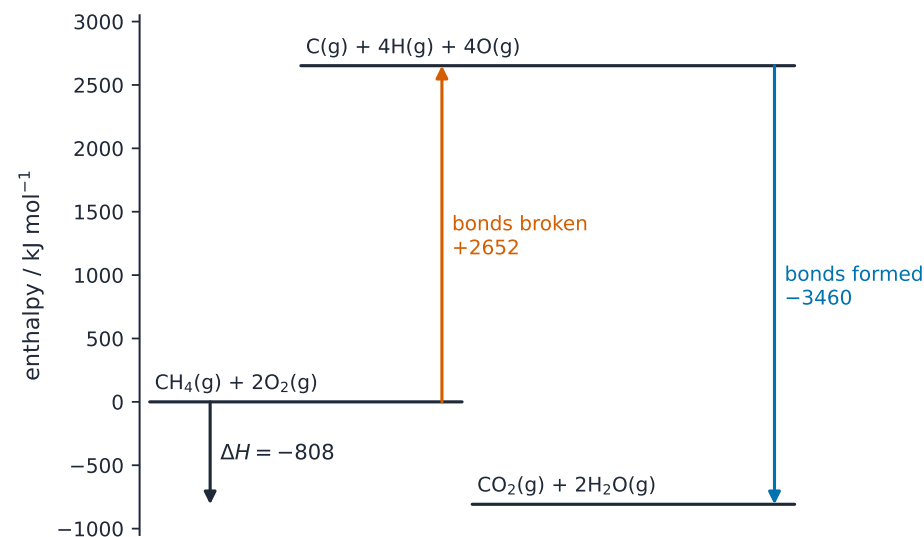
$$\Delta H = \sum(\text{bonds broken}) - \sum(\text{bonds formed})$$

Worked example: $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$.

Bonds broken	Energy / kJ	Bonds formed	Energy / kJ
4 C–H: 4×414	1656	2 C=O: 2×804	1608
2 O=O: 2×498	996	4 O–H: 4×463	1852
total	2652	total	3460

$\Delta H = 2652 - 3460 = -808 \text{ kJ mol}^{-1}$. More energy is released by forming bonds than

is needed to break them, so the reaction is exothermic.



Every bond is broken to gaseous atoms, then the new bonds form; ΔH is the difference

■ Why the answer is only approximate

The measured value for this reaction is -802 kJ mol^{-1} (with water as a gas). The bond enthalpy answer differs because:

- **average** bond enthalpies are used, not the values for these exact molecules;
- bond enthalpies apply only to **gases**. If a substance is a liquid or solid, the energy to change its state is missing. With liquid water as a product the true value is -890 kJ mol^{-1} .

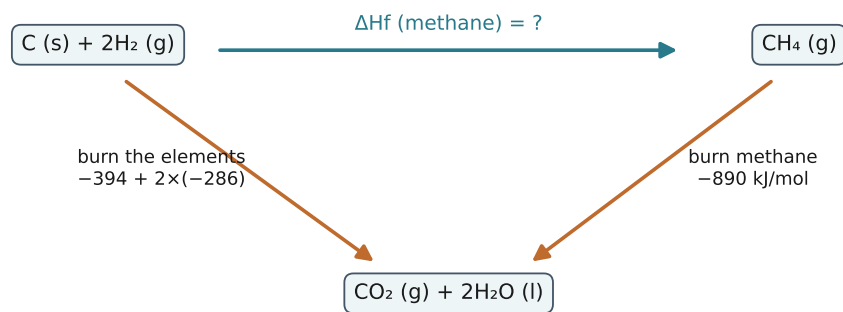
Hess's law

■ Same start, same end

Hess's law: the enthalpy change of a reaction is the same whatever route is taken, as long as the initial and final conditions are the same. Enthalpy is a **state function**: it depends only on where you start and where you finish, like the difference in height between the bottom and the top of a mountain, whichever path you climb.

This lets you calculate enthalpy changes that cannot be measured directly. Build an **energy cycle** with the target reaction on one route and reactions with known ΔH on the other.

Hess cycle: same start, same end → same total energy change



Same start, same end: $\Delta H_f = -394 + 2(-286) - (-890) = -76\text{ kJ/mol}$

■ Working with equations

You can also combine equations directly, like algebra:

- If you **reverse** an equation, change the sign of ΔH .
- If you **multiply** an equation by a number, multiply ΔH by the same number.
- Add the equations and their ΔH values; species that appear on both sides cancel.

Worked example: find ΔH for $\text{C(s)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{CO(g)}$, which cannot be measured because some CO_2 always forms.

- (1) $\text{C(s)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)}$, $\Delta H = -394\text{ kJ mol}^{-1}$
- (2) $\text{CO(g)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)}$, $\Delta H = -283\text{ kJ mol}^{-1}$

Keep (1) and reverse (2): $\text{CO}_2 \rightarrow \text{CO} + \frac{1}{2}\text{O}_2$, $\Delta H = +283$. Adding them, CO_2 cancels and $\frac{1}{2}\text{O}_2$ is left on the left:

$$\Delta H = -394 + 283 = -111\text{ kJ mol}^{-1}$$

■ Worked example: three equations

Find ΔH for $2\text{C(s)} + \text{H}_2\text{(g)} \rightarrow \text{C}_2\text{H}_2\text{(g)}$, given:

- (1) $\text{C(s)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)}$, $\Delta H = -394$
- (2) $\text{H}_2\text{(g)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{H}_2\text{O(l)}$, $\Delta H = -286$
- (3) $\text{C}_2\text{H}_2\text{(g)} + \frac{5}{2}\text{O}_2\text{(g)} \rightarrow 2\text{CO}_2\text{(g)} + \text{H}_2\text{O(l)}$, $\Delta H = -1301$

Take $2 \times (1)$ (two C on the left), $1 \times (2)$, and reverse (3) (C_2H_2 on the right):

$$\Delta H = 2(-394) + (-286) + 1301 = +227\text{ kJ mol}^{-1}$$

Check that the oxygen, CO_2 and water cancel: $\frac{5}{2}\text{O}_2$ on each side, 2CO_2 on each side, one H_2O on each side.

Practice

1. Explain, in terms of bonds, why the combustion of methane is exothermic.
2. Use bond enthalpies (H–H 436, Cl–Cl 242, H–Cl 431 kJ mol^{-1}) to calculate ΔH for $\text{H}_2\text{(g)} + \text{Cl}_2\text{(g)} \rightarrow 2\text{HCl(g)}$.
3. Use bond enthalpies (C=C 614, H–H 436, C–C 346, C–H 414 kJ mol^{-1}) to calculate ΔH for $\text{C}_2\text{H}_4\text{(g)} + \text{H}_2\text{(g)} \rightarrow \text{C}_2\text{H}_6\text{(g)}$.
4. Give two reasons why an enthalpy change calculated from average bond enthalpies differs from the measured value.
5. State Hess's law.
6. Given $\text{S(s)} + \text{O}_2\text{(g)} \rightarrow \text{SO}_2\text{(g)}$, $\Delta H = -297$, and $2\text{SO}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2\text{SO}_3\text{(g)}$, $\Delta H = -198\text{ kJ mol}^{-1}$, calculate ΔH for $\text{S(s)} + \frac{3}{2}\text{O}_2\text{(g)} \rightarrow \text{SO}_3\text{(g)}$.

Answers.

1. Bonds broken: 4 C–H and 2 O=O need 2652 kJ. Bonds formed: 2 C=O and 4 O–H release 3460 kJ. More energy is released forming the new bonds than is needed to break the old ones.
2. Broken: $436 + 242 = 678$. Formed: $2 \times 431 = 862$. $\Delta H = 678 - 862 = -184\text{ kJ mol}^{-1}$.
3. Broken: C=C 614 and H–H 436, total 1050 (the four C–H bonds are not broken). Formed: C–C 346 and two new C–H $2 \times 414 = 828$, total 1174. $\Delta H = 1050 - 1174 = -124\text{ kJ mol}^{-1}$.
4. Average bond enthalpies are means over many compounds, not the values for these molecules; bond enthalpies apply to gases, so any liquid or solid needs extra energy for the change of state.
5. The enthalpy change of a reaction is independent of the route taken, provided the initial and final conditions are the same.
6. Keep the first equation and halve the second: $\Delta H = -297 + \frac{1}{2}(-198) = -297 - 99 = -396\text{ kJ mol}^{-1}$.

Exam tips

- Bond breaking is endothermic and bond making exothermic; write “**broken minus formed**” before you substitute numbers.
- Count every bond: draw the structural formula, and leave out bonds that are not broken.
- Bond enthalpies are for **gases**; say so when you explain why a calculated value differs.
- When you reverse an equation, change the **sign**; when you multiply it, multiply ΔH .