IB DP CHEMISTRY

Reactivity 1: What Drives Chemical Reactions?

Reactivity 1.3 Energy from Fuels

Revision Notes · Standard and Higher Level

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Original notes prepared for the IB Diploma Programme Chemistry course (2025 syllabus)

What the syllabus requires

Reactivity 1.3 sits within the wider question of what makes reactions release energy. It builds on the enthalpy ideas of R1.1 and R1.2 and applies them to the fuels that power modern life. Use this checklist for a final revision sweep.

Understanding	You should be able to...
Combustion as a redox reaction	Write and balance equations for the complete and incomplete combustion of hydrocarbons and alcohols, and state the conditions and products of each.
Fossil fuels	Describe the origin and hydrocarbon composition of coal, oil and natural gas, and evaluate their advantages and limitations as energy sources.
Specific energy and energy density	Define both quantities with correct units, calculate them from data, and use them to compare fuels.
Environmental impact	Explain the roles of CO_2, CO, oxides of sulfur and nitrogen and particulates, and discuss carbon footprint qualitatively.
Biofuels	Describe the production of bioethanol and biodiesel and critically assess the claim that biofuels are 'carbon neutral'.
Fuel cells	Write the electrode half-equations for a hydrogen–oxygen fuel cell in acidic and alkaline media and contrast a fuel cell with combustion and with a battery.

> **Exam note:** The core of R1.3 is common to SL and HL. Higher Level candidates are expected to handle the electrode half-equations and efficiency comparisons in greater depth; those sections are flagged (HL) below.

1. Combustion of fuels

A **fuel** is a substance that releases useful energy when it burns. Combustion is the exothermic reaction of a fuel with oxygen; because the fuel is oxidised and oxygen reduced, every combustion is a redox reaction. Almost all common fuels are hydrocarbons (compounds of carbon and hydrogen only) or their oxygen-containing relatives such as alcohols.

Why combustion releases energy

Combustion is exothermic because the total energy released when new bonds form in the products (the strong C=O bonds in CO_2 and O−H bonds in H_2O) is greater than the energy absorbed to break the bonds in the fuel and in O_2. The enthalpy change of combustion, ΔH_c, is therefore negative, and can be estimated from bond enthalpies (R1.2) or measured by calorimetry (R1.1). The more energy released per gram or per mole, the better the fuel.

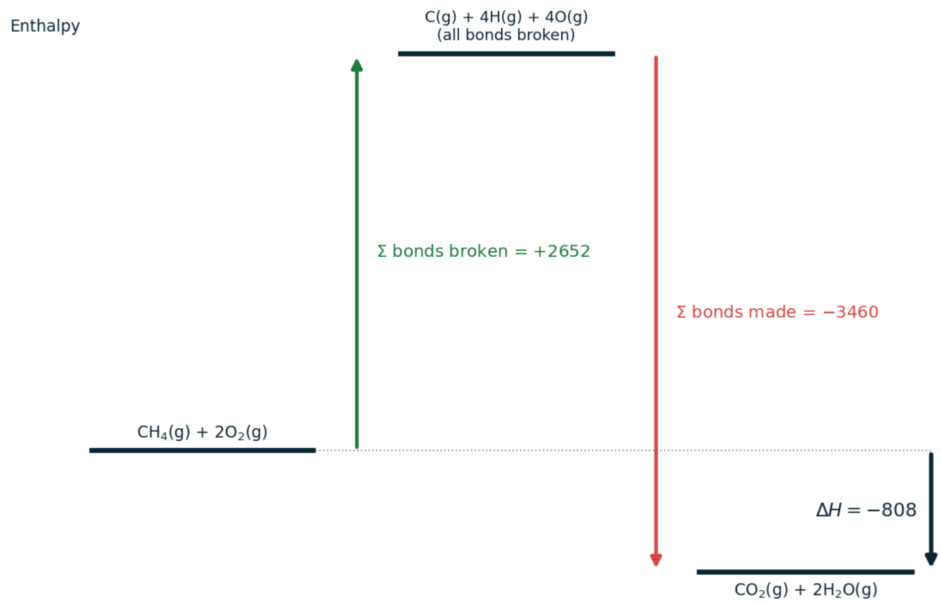

Figure 1. Estimating ΔH for the complete combustion of methane from average bond enthalpies: ΔH = bonds broken − bonds made = (+2652) − (+3460) = −808 kJ mol^{-1} (with gaseous water).

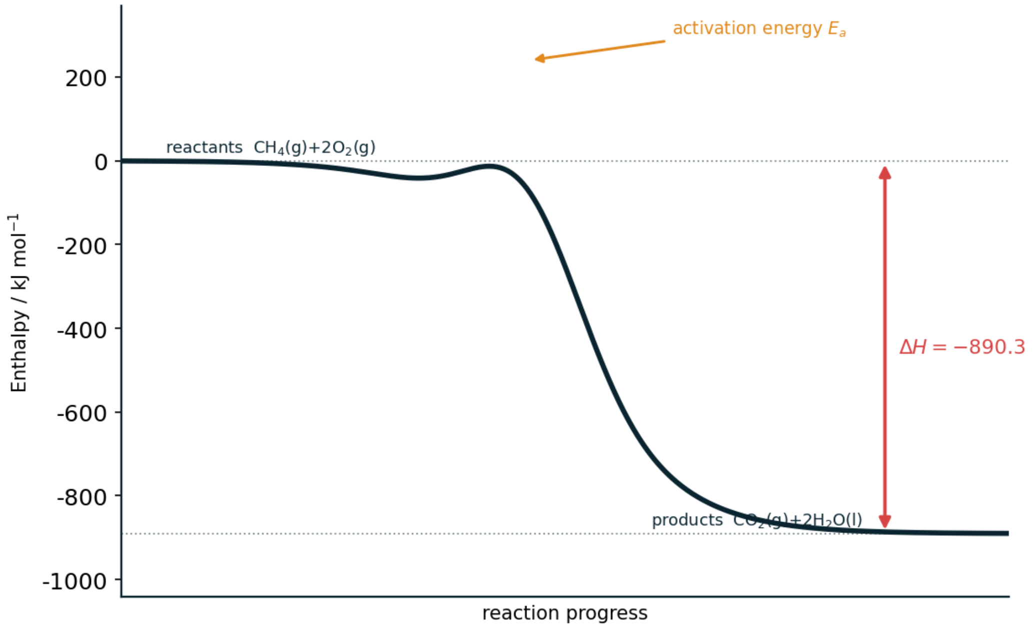

Figure 3. Enthalpy profile for the exothermic complete combustion of methane. Computed from standard formation enthalpies, ΔH = −890.3 kJ mol^{-1}; an activation energy is still needed to start it.

Complete combustion

When the oxygen supply is plentiful, combustion is **complete**: carbon is fully oxidised to carbon dioxide and hydrogen to water. This releases the maximum energy per mole of fuel and produces a clean, non-luminous blue flame.

methane + oxygen → carbon dioxide + water

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

For any hydrocarbon C_xH_y the general complete-combustion equation is:

$$C_xH_y + (x + y/4)\ O_2 \rightarrow x\ CO_2 + (y/2)\ H_2O$$

Incomplete combustion

When oxygen is limited, combustion is **incomplete**: there is not enough oxygen to convert all the carbon to CO_2. Carbon is only partially oxidised, giving a mixture that can include carbon monoxide (CO) and carbon itself as soot (C). The flame is yellow, luminous and sooty, and less energy is released per mole of fuel because the carbon is not fully oxidised.

$$2CH_4(g) + 3O_2(g) \rightarrow 2CO(g) + 4H_2O(l)\ \text{(limited oxygen)}$$

$$CH_4(g) + O_2(g) \rightarrow C(s) + 2H_2O(l)\ \text{(severely limited oxygen)}$$

Feature	Complete combustion	Incomplete combustion
Oxygen supply	Plentiful / excess	Limited / restricted
Carbon products	CO_2 only	CO and/or C (soot), often with some CO_2
Flame	Blue, non-luminous, hot	Yellow, luminous, sooty, cooler
Energy released	Maximum per mole of fuel	Less than complete
Hazard	CO_2 is a greenhouse gas	CO is a toxic gas; soot is a particulate

> **Exam technique:** Read the question for the phrase 'limited' or 'insufficient' oxygen. If you see it, the products must include CO or C — not just CO_2. Balance oxygen last, and remember fractional coefficients are allowed unless whole numbers are demanded.

2. Fossil fuels and comparing energy content

Fossil fuels formed over millions of years from the buried remains of once-living organisms, transformed under heat and pressure in the absence of air. They are concentrated stores of chemical energy but are finite (non-renewable) on any human timescale.

Fossil fuel	Origin	Composition	Notes
Coal	Compressed plant matter in ancient swamps	Mostly carbon, with some hydrogen, sulfur, nitrogen and mineral ash	Solid; highest CO_2 and SO_2 per unit energy
Crude oil (petroleum)	Marine micro-organisms buried in sediments	Complex mixture of liquid hydrocarbons (alkanes, cycloalkanes, arenes)	Refined by fractional distillation into useful fractions
Natural gas	Formed alongside oil	Mainly methane, CH_4, with some ethane and propane	Cleanest-burning fossil fuel; high H:C ratio

Advantages and limitations

- **Advantages:** high energy content, relatively cheap and abundant at present, established distribution networks, and easy to transport and store (especially liquids).
- **Limitations:** non-renewable and finite; combustion releases CO_2 (climate change) and pollutants (SO_x, NO_x, particulates, CO); extraction and transport carry environmental risks such as oil spills.

Renewable and non-renewable sources

It helps to place fuels in a wider map of energy resources. Non-renewable sources are consumed faster than they form; renewable sources are replenished on a human timescale.

Type	Examples	Key feature
Non-renewable	Coal, oil, natural gas, nuclear (uranium)	Finite reserves; fossil fuels release CO_2 on burning
Renewable	Biofuels, hydrogen (if from renewables), solar, wind, hydro	Replenished naturally; generally lower net CO_2

Specific energy and energy density

Two related measures let us compare how much energy a fuel packs. Do not confuse them.

Quantity	Definition	Typical unit
Specific energy	Energy released per unit **mass** of fuel	$kJ\ g^{-1}$ (or $MJ\ kg^{-1}$)
Energy density	Energy released per unit **volume** of fuel	$kJ\ dm^{-3}$ (or $MJ\ dm^{-3}$)

Specific energy matters where mass is critical (aircraft, rockets); energy density matters where space is critical (a car's fuel tank). Hydrogen has an outstanding specific energy but, as a gas, a very poor energy density — which is exactly why storing it is such a challenge (Section 5).

They are linked through density:

$$\textbf{energy density} = \textbf{specific energy} \times \textbf{density of fuel}$$

Typical values for some common fuels (approximate, for comparison) are shown below. Notice how ethanol's lower specific energy means a larger tank is needed for the same range, while hydrogen's high specific energy is undercut by its poor energy density as a gas.

Fuel	Specific energy / $kJ\ g^{-1}$	Comment
Hydrogen, H_2	142	Highest per gram; very low energy density as a gas
Natural gas (methane)	55	High H:C ratio, cleanest-burning fossil fuel
Petrol (gasoline)	47	Convenient liquid; high energy density
Diesel	45	Slightly denser than petrol
Ethanol (bioethanol)	30	Lower than petrol; larger tank for same range
Coal	24-33	Solid; high carbon, more CO_2 released per kJ

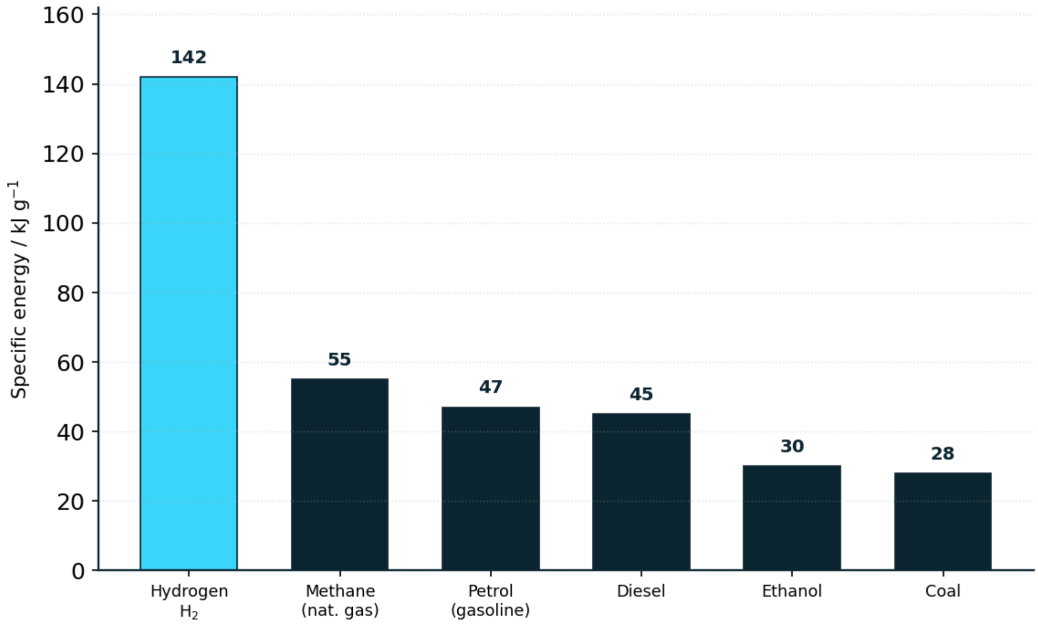

Figure 2. Specific energy (energy released per gram) of common fuels. Hydrogen (142 kJ g^{-1}) is about 4.7× the value for ethanol (30 kJ g^{-1}), yet as a gas it has a very low energy density.

Worked example 1 — balancing combustion equations

Write balanced equations for (a) the complete combustion of octane, C_8H_{18}, and (b) the complete combustion of ethanol, C_2H_5OH.

(a) Carbon: 8 CO_2. Hydrogen: 18 H → 9 H_2O. Oxygen needed = 8 × 2 + 9 = 25 atoms = 12.5 O_2. Multiply through by 2 for whole numbers:

$$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$$

(b) Ethanol already supplies one O atom. Carbon: 2 CO_2. Hydrogen: 3 H_2O. Oxygen in products = 4 + 3 = 7; the molecule provides 1, so O_2 supplies 6 atoms = 3 O_2:

$$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$$

Worked example 2 — specific energy

Burning 1.00 g of propane, C_3H_8, releases 50.4 kJ. Calculate (a) the specific energy, and (b) the molar enthalpy of combustion. ($M(C_3H_8)$ = 44.11 g mol^{-1})

(a) Specific energy = energy / mass = 50.4 kJ / 1.00 g = **50.4 kJ g^{-1}**.

(b) Per mole: ΔH_c = −50.4 × 44.11 = **−2220 kJ mol^{-1}** (negative because combustion is exothermic).

Worked example 3 — energy density and comparing fuels

A fuel has a specific energy of 34.2 kJ g^{-1} and a density of 0.79 g cm^{-3} (= 790 g dm^{-3}). Find its energy density and compare it with hydrogen gas (specific energy 142 kJ g^{-1}, density ≈ 0.090 g dm^{-3} at room conditions).

Fuel: energy density = 34.2 × 790 = **2.70×10^4 kJ dm^{-3}**.

Hydrogen: energy density = 142 × 0.090 ≈ **12.8 kJ dm^{-3}**.

So although hydrogen has a far higher specific energy, its energy density as a gas is about 2000 × smaller — a huge tank would be needed. This is the core storage problem for hydrogen as a vehicle fuel.

3. Environmental impact of burning fuels

Every combustion product has an environmental story. Examiners reward candidates who connect a specific pollutant to a specific effect and mechanism.

Product	Source	Environmental / health effect
Carbon dioxide, CO_2	Complete combustion of all carbon fuels	Greenhouse gas: absorbs and re-emits infrared radiation, enhancing the greenhouse effect and driving climate change
Carbon monoxide, CO	Incomplete combustion	Toxic: binds to haemoglobin (forming carboxyhaemoglobin) more strongly than O_2, reducing oxygen transport in the blood
Oxides of sulfur, SO_x	Sulfur impurities, mainly in coal and heavy oil	Form sulfurous / sulfuric acid in the air — acid deposition
Oxides of nitrogen, NO_x	N_2 and O_2 reacting at high engine temperatures	Acid deposition and photochemical smog
Particulates (soot, C)	Incomplete combustion, especially diesel	Respiratory damage; global dimming; deposits on surfaces

The greenhouse effect and climate change

Short-wavelength radiation from the Sun passes through the atmosphere and warms the Earth's surface. The surface re-emits energy as longer-wavelength infrared. Greenhouse gases — CO_2, methane (CH_4) and water vapour among them — absorb this infrared and re-radiate part of it back to the surface, raising the average temperature. Rising CO_2 from fossil-fuel combustion enhances this effect, contributing to global warming and climate change.

Acid deposition

Oxides of sulfur and nitrogen dissolve and react in atmospheric moisture to form acids (for example $SO_2 \rightarrow H_2SO_3$; further oxidation gives H_2SO_4), which fall as acid rain. Effects include acidification of lakes and soils, damage to forests, and corrosion of limestone buildings and metal structures.

Carbon footprint

The **carbon footprint** of an activity is the total amount of greenhouse gas it releases, usually expressed as a mass of CO_2 equivalent. Qualitatively, a fuel with a higher H:C ratio (such as natural gas) produces less CO_2 per unit of energy than one with a low H:C ratio (such as coal), so switching from coal to gas lowers the carbon footprint of electricity generation — although it remains a fossil fuel.

Worked example 5 - mass of CO_2 from combustion (carbon footprint)

Calculate the mass of CO_2 produced when 1.00 kg of methane, CH_4, undergoes complete combustion. ($M(CH_4) = 16.05\ g\ mol^{-1}$; $M(CO_2) = 44.01\ g\ mol^{-1}$)

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$, so 1 mol of CH_4 gives 1 mol of CO_2.

$n(CH_4) = 1000 / 16.05 = 62.3$ mol, therefore $n(CO_2) = 62.3$ mol.

$m(CO_2) = 62.3 \times 44.01 =$ **2.74×10^3 g** $\approx$ **2.74 kg**. Each kilogram of methane burned adds about 2.7 kg of CO_2 to the atmosphere - the basis of a quantitative carbon-footprint estimate.

Common misconception: CO_2 and CO are not interchangeable. CO_2 is the greenhouse-gas product of *complete* combustion; CO is the *toxic* product of *incomplete* combustion. Attributing CO_2 toxicity or CO climate-forcing loses marks.

4. Biofuels

Biofuels are made from recently living biomass rather than from long-buried fossil deposits, so they are considered **renewable**. The two most examined are bioethanol and biodiesel.

Bioethanol

Bioethanol is ethanol made by the **fermentation** of sugars from crops such as sugar cane or maize, using yeast enzymes in the absence of air:

$$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g) \text{ (fermentation)}$$

The ethanol is concentrated by distillation and burned, alone or blended with petrol (for example E10):

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l) \text{ (combustion)}$$

Biodiesel

Biodiesel is made from vegetable oils or animal fats. In **transesterification**, the large triglyceride molecules are reacted with a short-chain alcohol (usually methanol) in the presence of a base catalyst, replacing the glycerol backbone and producing smaller fatty-acid methyl ester molecules — biodiesel — plus glycerol as a by-product. These esters have a viscosity and volatility close to those of conventional diesel, so they can be used in existing engines.

triglyceride + 3 methanol → 3 methyl esters (biodiesel) + glycerol

The 'carbon-neutral' argument and its limits

Biofuels are often called **carbon neutral**: the CO_2 released when they burn was recently absorbed from the atmosphere by the crop during photosynthesis, so in principle there is no net addition of CO_2. In practice this is only approximately true, because:

- fossil fuels are usually burned to plant, harvest, process and transport the crop;
- fertiliser manufacture and land clearing (releasing stored carbon) add emissions;
- using farmland for fuel can displace food production and drive deforestation.

So biofuels reduce net CO_2 compared with fossil fuels, but full carbon neutrality is an idealisation rather than a strict fact.

	Advantages	Disadvantages
Biofuels vs fossil fuels	Renewable; approximately carbon neutral; biodegradable; usually lower SO_x; can use existing engines/infrastructure	Lower specific energy than diesel/petrol; competes with food crops and land; energy input to grow and process; seasonal supply

Newer 'second-generation' biofuels are made from non-food biomass such as agricultural waste, wood residues or algae. These ease the food-versus-fuel conflict and can use land unsuitable for food crops, but the technology to break down tough cellulose economically is still developing.

> **Exam technique:** If asked to 'evaluate', give both sides and a judgement. State the carbon-neutral claim, then qualify it with at least one reason it is not perfectly true — that nuance is what distinguishes top answers.

5. Fuel cells

A **fuel cell** converts the chemical energy of a fuel directly into electrical energy through controlled redox half-reactions at two electrodes, without burning the fuel. The most important example is the hydrogen–oxygen fuel cell, in which the only product is water.

Component	Role
Fuel inlet (anode side)	Supplies hydrogen continuously to the negative electrode
Oxidant inlet (cathode side)	Supplies oxygen (or air) to the positive electrode
Porous electrodes + catalyst	Provide a large surface area; catalyst (e.g. platinum) speeds the electron-transfer half-reactions
Electrolyte / membrane	Conducts ions (H^+ or OH^-) internally but forces electrons through the external circuit
External circuit	Carries electrons from anode to cathode, delivering electrical work to the load

Hydrogen is oxidised at the negative electrode (anode); oxygen is reduced at the positive electrode (cathode). The half-equations depend on the electrolyte.

In acidic electrolyte

$$\textbf{Anode (oxidation): } 2H_2(g) \rightarrow 4H^+(aq) + 4e^-$$

$$\textbf{Cathode (reduction): } O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$$

In alkaline electrolyte

$$\textbf{Anode (oxidation): } 2H_2(g) + 4OH^-(aq) \rightarrow 4H_2O(l) + 4e^-$$

$$\textbf{Cathode (reduction): } O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$$

Overall reaction (either medium):

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

How a fuel cell differs from combustion and from a battery

- **vs combustion:** the same overall reaction occurs, but a fuel cell separates oxidation and reduction at two electrodes so that electrons flow through an external circuit, delivering electrical energy directly instead of heat. This avoids the efficiency limits of a heat engine.
- **vs battery:** a battery contains a fixed amount of stored reactants and runs down; a fuel cell has its reactants supplied continuously from outside and keeps producing electricity as long as fuel and oxygen are fed to it.

Advantages and challenges

- **Advantages:** high efficiency, quiet operation, and only water as the product at the point of use — no CO_2, CO, NO_x or particulates.
- **Challenges:** hydrogen is difficult to store safely (very low energy density as a gas; high-pressure or cryogenic tanks) and most hydrogen is currently produced from fossil fuels (steam reforming of methane), which shifts rather than removes the emissions. Electrolysis using renewable electricity gives 'green' hydrogen but is more expensive.

Context — direct methanol fuel cell

Some fuel cells run on liquid methanol instead of hydrogen. Methanol is far easier to store and transport than hydrogen gas, but its combustion product stream includes CO_2, so it is not emission-free at the point of use. This illustrates the recurring trade-off between ease of storage and cleanliness.

> **Common misconception:** A hydrogen fuel cell is only as clean as its hydrogen supply. 'Zero emission' refers to the point of use; if the hydrogen came from steam reforming of methane, CO_2 was released upstream during its manufacture.

6. (HL) Electrode half-reactions and efficiency in more depth

(HL) At Higher Level you should construct and balance the electrode half-equations confidently and reason about the energy advantage of a fuel cell.

Building a half-equation from scratch (HL)

To balance a half-equation in acidic solution: (1) balance the atom being oxidised or reduced; (2) balance oxygen with H_2O; (3) balance hydrogen with H^+; (4) balance charge with electrons. In alkaline solution, add OH^- to both sides to neutralise any H^+, combining them into water. Applying this to O_2 reduction in acid gives $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$, exactly the cathode half-equation above.

Why fuel cells beat combustion engines (HL)

A petrol or diesel engine is a **heat engine**: it first turns chemical energy into heat, then heat into work, so its efficiency is capped by thermodynamics (typically around 25–35% of the fuel energy reaches the wheels). A fuel cell converts chemical energy to electrical energy directly, without an intermediate heat step, so practical efficiencies of 40–60% are achievable. Energy efficiency here means:

$$\text{efficiency} = (\text{useful energy output} / \text{total energy input}) \times 100\%$$

Worked example 4 (HL) — fuel-cell half-equations and efficiency

(a) Write the anode and cathode half-equations for a hydrogen–oxygen fuel cell operating in alkaline (KOH) electrolyte and show they combine to the overall equation. (b) If a fuel cell delivers 58% of the fuel's energy as electricity while a combustion engine delivers 30%, how many times more of the fuel's energy is usefully captured by the cell?

(a) Anode: $2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$. Cathode: $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$.

Adding: the $4e^-$, the OH^- and 2 of the H_2O cancel, leaving $2H_2 + O_2 \rightarrow 2H_2O$.

(b) Ratio = 58 / 30 = **1.9**, so the fuel cell captures about 1.9 × as much of the fuel's energy usefully.

HL note: When you combine half-equations, make the electrons equal first, then add and cancel species that appear identically on both sides (electrons, and any excess H_2O or OH^-). The electrons must always cancel completely.

7. Nuclear energy — a brief comparison

For balance, the syllabus places chemical fuels alongside nuclear energy sources, which release energy from changes in the nucleus rather than from chemical bonds. Only a qualitative appreciation is needed.

Process	What happens	Points to note
Nuclear fission	A heavy nucleus (e.g. uranium-235) splits into smaller nuclei when it absorbs a neutron, releasing large energy	Used in current power stations; produces radioactive waste; no CO_2 at generation
Nuclear fusion	Light nuclei (e.g. isotopes of hydrogen) join to form a heavier nucleus, releasing very large energy	Powers the Sun; needs extreme temperatures; not yet a practical large-scale source on Earth

Both release far more energy per unit mass than any chemical fuel and emit no CO_2 at the point of generation, but each carries its own challenges — radioactive waste and safety for fission, and the enormous engineering demands of sustained fusion.

8. Common pitfalls

- Writing CO_2 as a product when the question specifies limited oxygen — the products should be CO and/or C.
- Confusing specific energy (per unit mass, kJ g^{-1}) with energy density (per unit volume, kJ dm^{-3}).
- Claiming biofuels are perfectly carbon neutral without noting the fossil-fuel inputs to grow and process them.
- Muddling CO_2 (greenhouse gas, complete combustion) with CO (toxic gas, incomplete combustion).
- Forgetting to balance charge with electrons in fuel-cell half-equations, or leaving electrons uncancelled in the overall equation.
- Calling a hydrogen fuel cell 'zero emission' overall when the hydrogen was made from methane.

- Dropping the DejaVu minus sign or the negative sign on an exothermic ΔH_c value.

9. Quick reference

Item	Statement
Complete combustion (general)	$C_xH_y + (x + y/4)O_2 \rightarrow x\ CO_2 + (y/2)H_2O$
Incomplete combustion	Products include CO and/or C (soot) when oxygen is limited
Specific energy	Energy per unit mass; unit kJ g^{-1}
Energy density	Energy per unit volume; unit kJ dm^{-3}; = specific energy × density
Fermentation	$C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$
Fuel cell (acid)	Anode $2H_2 \rightarrow 4H^+ + 4e^-$; Cathode $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$
Fuel cell (alkaline)	Anode $2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$; Cathode $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$
Fuel cell overall	$2H_2 + O_2 \rightarrow 2H_2O$
Pollutants to link	CO_2: greenhouse; CO: toxic; SO_x/NO_x: acid rain; particulates: respiratory harm

10. Test yourself

Attempt these without notes; full worked answers follow.

1. Write balanced equations for the complete combustion of (a) propane, C_3H_8, and (b) butan-1-ol, C_4H_9OH.

2. Write an equation for the incomplete combustion of ethane, C_2H_6, that gives carbon monoxide and water only.

3. Burning 2.50 g of methanol releases 56.8 kJ. Find its specific energy and its molar enthalpy of combustion. ($M(CH_3OH)$ = 32.05 g mol^{-1})

4. Fuel X has specific energy 43.0 kJ g^{-1} and density 0.72 g cm^{-3}. Fuel Y has specific energy 27.0 kJ g^{-1} and density 0.79 g cm^{-3}. Which has the greater energy density?

5. Explain why carbon monoxide is toxic and why it forms during combustion.

6. Discuss whether bioethanol can properly be described as carbon neutral.

7. Write the anode and cathode half-equations for a hydrogen–oxygen fuel cell in acidic solution, and give the overall equation.

8. Give one advantage and one challenge of using hydrogen fuel cells in cars compared with a petrol engine.

9. Explain, in terms of H:C ratio, why burning natural gas has a smaller carbon footprint per unit of energy than burning coal.

10. A car tank holds 45 dm^3 of petrol (energy density 3.4×10^4 kJ dm^{-3}). How much energy is stored, and why would a bioethanol tank of the same volume store less?

Answers

1. (a) $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$. (b) Butanol supplies 1 O; products $4CO_2 + 5H_2O$ need 13 O atoms, minus 1 from the molecule = 12 from O_2 = 6 O_2: $C_4H_9OH + 6O_2 \rightarrow 4CO_2 + 5H_2O$.

2. $2C_2H_6 + 5O_2 \rightarrow 4CO(g) + 6H_2O$. (Carbon → CO; oxygen atoms on the right = 4 + 6 = 10 = 5 O_2.)

3. Specific energy = 56.8 / 2.50 = **22.7 kJ g^{-1}**. Per mole: $\Delta H_c = -22.7 \times 32.05 \approx$ **−728 kJ mol^{-1}**.

4. X: $43.0 \times 720 = 3.10 \times 10^4$ kJ dm^{-3}. Y: $27.0 \times 790 = 2.13 \times 10^4$ kJ dm^{-3}. **Fuel X** has the greater energy density, despite the smaller density, because of its higher specific energy.

5. CO forms by incomplete combustion when oxygen is limited, so carbon is only partly oxidised. It is toxic because it binds strongly to haemoglobin (as carboxyhaemoglobin), blocking the sites that normally carry oxygen and so reducing oxygen delivery to tissues.

6. The CO_2 released on burning was recently taken from the air by the crop during photosynthesis, so in principle there is no net CO_2 addition. However fossil fuels are used to grow, harvest, distil and transport it, and land-use change adds emissions, so it is better described as approximately, not perfectly, carbon neutral.

7. Anode: $2H_2 \rightarrow 4H^+ + 4e^-$. Cathode: $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$. Overall: $2H_2 + O_2 \rightarrow 2H_2O$.

8. Advantage: the only product at the point of use is water — no CO_2, CO, NO_x or particulates — and the cell is more efficient than a heat engine. Challenge: hydrogen is hard to store safely (low energy density as a gas) and is often made from fossil fuels, shifting the emissions upstream.

9. Natural gas (mainly CH_4) has a high H:C ratio, so much of its energy comes from oxidising hydrogen to water rather than carbon to CO_2. Coal is almost all carbon (low H:C), so it releases more CO_2 for the same amount of energy. Hence switching from coal to gas lowers the carbon footprint of generation, though gas is still a fossil fuel.

10. Energy = energy density × volume = $3.4 \times 10^4 \times 45$ = **1.5×10^6 kJ**. A bioethanol tank of the same volume stores less because ethanol has a lower specific energy (and lower energy density) than petrol, so the same volume holds less usable energy - giving a shorter driving range.