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IB DIPLOMA CHEMISTRY · FIRST ASSESSMENT 2025

Reactivity 1.3 Energy from Fuels

Booklet 4 · Understandings 1.3.1 – 1.3.5

Complete and incomplete combustion, fossil fuels and the greenhouse effect, biofuels, and fuel cells that turn chemical energy directly into electrical energy

Complete revision notes, 14 worked examples, model answers & 31 practice questions · Standard Level & Higher Level

How the level tags work in this booklet

SL + HL Every understanding in Reactivity 1.3 is common to both levels, so the whole booklet is core material.

HL One linking box (why CO_2 is a greenhouse gas, from HL Structure 3.2) and the ΔH_f calculations in section 3 use HL skills from Booklet 3. ★ marks harder, multi-step questions.

Inside this booklet

1. What the syllabus actually asks of you
2. 1.3.1 — Combustion reactions
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1 What the syllabus actually asks of you

Reactivity 1.1 and 1.2 were about measuring and calculating enthalpy changes. Reactivity 1.3 applies those ideas to the question that matters most outside the laboratory: how do we get the energy we need, and at what cost? Burning fuels still provides most of the world's primary energy. The topic asks you to write equations for complete and incomplete combustion, compare fossil fuels and biofuels using numbers, explain the link between carbon dioxide and the greenhouse effect, and describe how a fuel cell produces electricity without combustion. Expect equation-writing and mole calculations in Paper 1, and "evaluate", "compare" and "explain" questions in Paper 2.

Understanding	What it says	What you must be able to do	Level
1.3.1 Combustion	Reactive metals, non-metals and organic compounds undergo combustion reactions when heated in oxygen.	Deduce equations for reactions of combustion, including hydrocarbons and alcohols.	SL+HL
1.3.2 Incomplete combustion	Incomplete combustion of organic compounds, especially hydrocarbons, leads to the production of carbon monoxide and carbon.	Deduce equations for the incomplete combustion of hydrocarbons and alcohols.	SL+HL
1.3.3 Fossil fuels	Fossil fuels include coal, crude oil and natural gas, which have different advantages and disadvantages.	Evaluate the amount of carbon dioxide added to the atmosphere when different fuels burn. Understand the link between carbon dioxide levels and the greenhouse effect. The tendency for incomplete combustion and energy released per unit mass should be covered.	SL+HL
1.3.4 Biofuels	Biofuels are produced from the biological fixation of carbon over a short period of time through photosynthesis.	Understand the difference between renewable and non-renewable energy sources. Consider the advantages and disadvantages of biofuels. The reactants and products of photosynthesis should be known.	SL+HL
1.3.5 Fuel cells	A fuel cell can be used to convert chemical energy from a fuel directly to electrical energy.	Deduce half-equations for the electrode reactions in a fuel cell. Hydrogen and methanol should be covered as fuels for fuel cells. **The use of proton exchange membranes will not be assessed.**	SL+HL

Linking questions and connections for 1.3.1 – 1.3.5

- **Reactivity 2.2:** why is high activation energy often considered to be a useful property of a fuel? → section 2.4
- **Reactivity 3.2:** which species are the oxidizing and reducing agents in a combustion reaction? What are the main differences between a fuel cell and a primary (voltaic) cell? → sections 2.4 and 8
- **Reactivity 2.1:** how does limiting the supply of oxygen in combustion affect the products and increase health risks? → section 3
- **Structure 3.2:** why do larger hydrocarbons have a greater tendency to undergo incomplete combustion? → section 3.3. HL Why is carbon dioxide described as a greenhouse gas? → section 6.3
- **Structure 1.4:** combustion analysis gives the empirical formula of a compound. → section 2.5
- **Booklets 1–3:** ΔH_c values, specific energy and ΔH_f calculations are used throughout.

2 Reactivity 1.3.1 — Combustion reactions SL + HL

2.1 What combustion is

Combustion is a reaction in which an element or compound burns in oxygen, releasing energy as heat and usually light. It was once thought that burning released a fire-like substance, "phlogiston". Lavoisier's careful weighing in the 1780s showed that substances actually **gain** mass when they burn, because they combine with oxygen from the air. This is a good example of a theory being replaced when it can no longer explain quantitative evidence.

2.2 Combustion of metals and non-metals

Reactive metals and many non-metals burn when heated in oxygen to form oxides. Recall from Structure 3.1 that **metal oxides are basic** (ionic) and **non-metal oxides are acidic** (covalent).

Element	Equation	Observation	Oxide
magnesium	$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$	intense white flame, white powder	basic, ionic
sodium	$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$	yellow-orange flame, white solid	basic, ionic
sulfur	$S(s) + O_2(g) \rightarrow SO_2(g)$	blue flame, choking toxic gas	acidic, covalent
carbon	$C(s) + O_2(g) \rightarrow CO_2(g)$	glows red; colourless gas	acidic, covalent
hydrogen	$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$	pale blue flame, "pop" in small amounts	neutral
phosphorus (white, P_4)	$P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$	ignites spontaneously in air; white smoke	acidic, covalent

Some non-metals form more than one oxide. With a limited supply of oxygen, phosphorus forms P_4O_6; magnesium burning in air also reacts with nitrogen to form Mg_3N_2.

2.3 Complete combustion of hydrocarbons and alcohols

Hydrocarbons and alcohols are used as fuels because they release a lot of energy at a useful rate. With a plentiful supply of oxygen, **complete combustion** breaks the carbon chain and fully oxidizes the compound: every carbon atom ends up in CO_2 and every hydrogen atom in H_2O.

Method — balancing combustion equations

1. Write the formula of the fuel and the products CO_2 and H_2O, with O_2 on the left.
2. Balance **C** first, then **H**.
3. Count the O atoms on the right. Subtract any O atoms in the fuel. Halve what is left: that is the coefficient of O_2.
4. Fractions are fine (and required for an enthalpy of combustion). Double everything if whole numbers are asked for.

Worked example 1 — balancing step by step

Fuel	Balance C and H	Balance O
pentane, C_5H_{12}	$C_5H_{12} + _O_2 \rightarrow 5CO_2 + 6H_2O$	10 + 6 = 16 O on the right → $8O_2$ **$C_5H_{12}(l) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(l)$**
ethanol, C_2H_5OH	$C_2H_5OH + _O_2 \rightarrow 2CO_2 + 3H_2O$	4 + 3 = 7 O; ethanol already has 1 → 6 O needed → $3O_2$ **$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$**
octane, C_8H_{18}	$C_8H_{18} + _O_2 \rightarrow 8CO_2 + 9H_2O$	16 + 9 = 25 O → $12½O_2$; doubled: **$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$**
propan-2-ol, C_3H_7OH	$C_3H_7OH + _O_2 \rightarrow 3CO_2 + 4H_2O$	6 + 4 = 10 O; minus 1 = 9 → $4½O_2$ **$C_3H_7OH(l) + 4½O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$**

Worked example 2 — general equations

For any hydrocarbon C_xH_y, there are x C atoms → x CO_2, and y H atoms → y/2 H_2O. Oxygen atoms needed = 2x + y/2, so O_2 = x + y/4:

$$C_xH_y + (x + {}^{y}/_{4})O_2 \rightarrow xCO_2 + {}^{y}/_{2}H_2O$$

For an alkane C_nH_{2n+2}: $C_nH_{2n+2} + {}^{(3n+1)}/_{2}O_2 \rightarrow nCO_2 + (n+1)H_2O$. For an alcohol $C_nH_{2n+1}OH$: $C_nH_{2n+1}OH + {}^{3n}/_{2}O_2 \rightarrow nCO_2 + (n+1)H_2O$.

Check: butane, n = 4, needs $(3 \times 4 + 1)/2 = 6½$ O_2; butan-1-ol needs $3 \times 4/2 = 6$ O_2.

2.4 Combustion as redox, and why fuels do not burn by themselves

All combustion reactions are **redox** reactions (Reactivity 3.2). Oxygen is the **oxidizing agent**: it gains electrons and its oxidation state falls from 0 to −2. The fuel is the **reducing agent**: in $Mg \rightarrow Mg^{2+}$ the metal loses electrons, and in $CH_4 \rightarrow CO_2$ the oxidation state of carbon rises from −4 to +4.

Fuels such as methane, petrol and wood are energetically unstable relative to their combustion products, yet they can be stored safely in air because the combustion reaction has a **high activation energy**. They are **kinetically stable**. A spark or flame supplies the activation energy; once started, the heat released keeps the reaction going. A high activation energy is therefore a *useful* property in a fuel: it allows safe storage and transport, and burning can be controlled.

2.5 Calculations from combustion equations

Worked example 3 — masses of products and reactants

Q. Calculate (a) the mass of carbon dioxide produced and (b) the mass of oxygen used when 2.90 g of butane, C_4H_{10} ($M = 58.14$ g mol^{-1}), undergoes complete combustion.

$C_4H_{10}(g) + 6½O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$

$n(C_4H_{10}) = 2.90 / 58.14 = 0.0499$ mol

(a) $n(CO_2) = 4 \times 0.0499 = 0.200$ mol; $m = 0.200 \times 44.01 =$ **8.78 g**

(b) $n(O_2) = 6.5 \times 0.0499 = 0.324$ mol; $m = 0.324 \times 32.00 =$ **10.4 g**

Notice that the products weigh more than the fuel. The extra mass is the oxygen from the air: Lavoisier's observation.

Worked example 4 — combustion analysis (link to Structure 1.4)

Q. 0.250 g of a compound containing only C, H and O burned completely to give 0.595 g CO_2 and 0.304 g H_2O. Determine its empirical formula.

mass C = 0.595 × 12.01/44.01 = 0.1624 g → n(C) = 0.01352 mol

mass H = 0.304 × 2.02/18.02 = 0.0341 g → n(H) = 0.0337 mol

mass O = 0.250 − 0.1624 − 0.0341 = 0.0535 g → n(O) = 0.00335 mol

ratio C : H : O = 0.01352 : 0.0337 : 0.00335 = 4.04 : 10.1 : 1 → **$C_4H_{10}O$**

All the carbon ends up in the CO_2 and all the hydrogen in the H_2O; oxygen is found by difference, because some of the oxygen in the products came from the air.

Examiner traps in 1.3.1

- Forgetting the O atom already in an alcohol when balancing O_2.
- Writing O instead of O_2, or P instead of P_4 for white phosphorus.
- Calling rusting "combustion". Rusting is oxidation, but there is no flame and it is very slow.
- "Combustion is always exothermic": nearly always, but $N_2 + O_2 \rightarrow 2NO$ (in hot engines and lightning) is endothermic.

3 Reactivity 1.3.2 — Incomplete combustion SL + HL

3.1 Carbon monoxide and soot

When the supply of oxygen is limited, or the fuel has a high percentage of carbon, there is not enough oxygen to oxidize every carbon atom fully. **Incomplete combustion** then produces **carbon monoxide**, CO, or **carbon** (soot), C, as well as water. The hydrogen is still oxidized to water because it reacts more readily than carbon.

Fuel and product	Equation
methane → carbon monoxide	$CH_4(g) + 1½O_2(g) \rightarrow CO(g) + 2H_2O(l)$ or $2CH_4 + 3O_2 \rightarrow 2CO + 4H_2O$
methane → carbon	$CH_4(g) + O_2(g) \rightarrow C(s) + 2H_2O(l)$
propane → carbon monoxide	$2C_3H_8(g) + 7O_2(g) \rightarrow 6CO(g) + 8H_2O(l)$
propane → carbon	$C_3H_8(g) + 2O_2(g) \rightarrow 3C(s) + 4H_2O(l)$
ethanol → carbon monoxide	$C_2H_5OH(l) + 2O_2(g) \rightarrow 2CO(g) + 3H_2O(l)$
general, to CO	$C_xH_y + (^x/_2 + ^y/_4)O_2 \rightarrow xCO + ^y/_2H_2O$

In practice a mixture of CO_2, CO and C forms, and a question may specify a ratio. A Bunsen burner shows the difference clearly: with the air hole open, methane burns with a hot, blue, non-luminous flame (complete); with the air hole closed, it gives a cooler, yellow, sooty flame (incomplete) because glowing carbon particles emit yellow light.

Worked example 5 — incomplete combustion equations

Q. (a) Write an equation for the incomplete combustion of octane in a car engine to form carbon monoxide and steam. (b) Write an equation for the combustion of hexane that produces equal amounts of CO_2 and CO.

(a) $C_8H_{18}(l) + _O_2(g) \rightarrow 8CO(g) + 9H_2O(g)$: 8 + 9 = 17 O → $8½O_2$

$C_8H_{18}(l) + 8½O_2(g) \rightarrow 8CO(g) + 9H_2O(g)$ (or $2C_8H_{18} + 17O_2 \rightarrow 16CO + 18H_2O$)

(b) 6 C atoms → $3CO_2$ + 3CO; 14 H → $7H_2O$; O atoms = 6 + 3 + 7 = 16 → $8O_2$

$C_6H_{14}(l) + 8O_2(g) \rightarrow 3CO_2(g) + 3CO(g) + 7H_2O(l)$

3.2 Incomplete combustion releases less energy

CO and C are only partly oxidized, so less energy is released than in complete combustion. This can be shown with standard enthalpies of formation (HL skill, Booklet 3), using ΔH_f: $C_6H_{14}(l)$ −199, $CO_2(g)$ −394, $CO(g)$ −111, $H_2O(l)$ −286 kJ mol^{-1}.

Worked example 6 — complete and incomplete combustion of hexane

complete: $C_6H_{14} + 9½O_2 \rightarrow 6CO_2 + 7H_2O$: $\Delta H = [6(-394) + 7(-286)] - (-199) =$ **−4167 kJ mol^{-1}**

to CO: $C_6H_{14} + 6½O_2 \rightarrow 6CO + 7H_2O$: $\Delta H = [6(-111) + 7(-286)] - (-199) =$ **−2469 kJ mol^{-1}**

to C: $C_6H_{14} + 3½O_2 \rightarrow 6C + 7H_2O$: $\Delta H = [0 + 7(-286)] - (-199) =$ **−1803 kJ mol^{-1}**

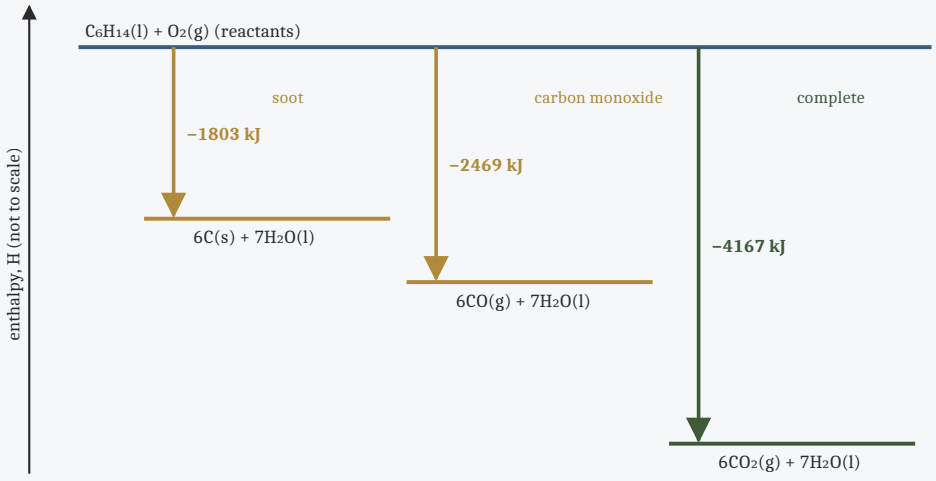


Figure 1 — The less completely the carbon is oxidized, the less energy is released. Forming CO wastes about 40 % of the available energy; forming soot wastes almost 57 %.

3.3 Which fuels burn incompletely most easily?

The higher the **percentage of carbon by mass**, the more oxygen is needed per gram of fuel and the greater the tendency for incomplete combustion and a sooty flame. Larger hydrocarbons also have stronger London (dispersion) forces (Structure 3.2), so they are less volatile and mix less easily with air in the gas phase; this also favours incomplete combustion.

Worked example 7 — predicting the sootiest flame

Compound	ethanol, C_2H_5OH	methane, CH_4	hexane, C_6H_{14}	cyclohexene, C_6H_{10}	methylbenzene, C_7H_8
M / g mol^{-1}	46.08	16.05	86.20	82.16	92.15
% C by mass	52.1	74.8	83.6	87.7	**91.2**
flame	clean, blue	clean	slightly sooty	sooty	very sooty, yellow

Aromatic compounds such as benzene and methylbenzene have the highest carbon content and give very smoky flames. Observing how sooty a flame is can help to distinguish organic compounds.

3.4 Why incomplete combustion matters: pollution and health

Pollutant	Source	Effect
carbon monoxide, CO	incomplete combustion in engines, faulty gas boilers and fires	colourless, odourless and toxic. It binds to **haemoglobin** in red blood cells much more strongly than oxygen (Hb + CO ⇌ HbCO), so less oxygen is carried round the body: headaches and dizziness at low levels, unconsciousness and death at high levels. Detectors are used in homes.
particulates (soot, C)	incomplete combustion of diesel, coal, wood	tiny particles (about 0.001–10 µm) are inhaled deep into the lungs, causing respiratory disease and cancer; they also blacken buildings and can catalyse other pollution reactions
unburnt hydrocarbons	engines	react with nitrogen oxides in sunlight to form photochemical smog
nitrogen oxides, NO and NO_2	$N_2 + O_2 \rightarrow 2NO$ at the high temperature inside engines; $2NO + O_2 \rightarrow 2NO_2$	acid rain, smog and breathing problems (not caused by incomplete combustion, but formed alongside it)

Catalytic converters in car exhausts reduce these emissions: $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$ oxidizes carbon monoxide, unburnt hydrocarbons are oxidized to CO_2 and H_2O, and a three-way converter also reduces nitrogen oxides: $2CO(g) + 2NO(g) \rightarrow 2CO_2(g) + N_2(g)$.

> **Model answer: "Explain why the incomplete combustion of fuels is undesirable." (3 marks)**
>
> Less energy is released per mole of fuel, because carbon is only partly oxidized to CO or C **[1]**. Carbon monoxide is toxic: it binds to haemoglobin and reduces the blood's ability to carry oxygen **[1]**. Soot particulates cause respiratory problems (and blacken surfaces/block burners) **[1]**.

Examiner traps in 1.3.2

- Producing CO and C but forgetting the water, or turning the hydrogen into H_2.
- Saying CO "causes global warming" or "acid rain". Its key hazard is toxicity.
- Saying "not enough fuel" instead of "not enough oxygen".
- Stating that CO is toxic without saying how: it combines with haemoglobin.

4 Reactivity 1.3.3 — Fossil fuels SL + HL

4.1 What makes a good fuel

An ideal fuel releases a large amount of energy, at a reasonable rate (not too fast, not too slow), is cheap, plentiful and easy to store and transport, and produces minimal pollution. No real fuel meets every criterion, so the choice always involves trade-offs.

4.2 Formation and the three fossil fuels

Fossil fuels are "sunshine in solid, liquid and gaseous form": their energy came from sunlight trapped by photosynthesis millions of years ago. They formed by the slow, partial decomposition of plant and animal matter buried **in the absence of air**, under high temperature and pressure. Oxygen was lost from the biological molecules faster than other elements, leaving **reduced** compounds that are mainly hydrocarbons or carbon. Because they take millions of years to form and are used far faster than they could be replaced, fossil fuels are **non-renewable**.

- **Coal** — a combustible sedimentary rock formed from plant remains. Heat and pressure increase its carbon content in stages; most coal is 80–90 % carbon by mass, and anthracite is over 90 %. It is the most abundant fossil fuel.
- **Crude oil** — a complex mixture of straight-chain and branched alkanes, cycloalkanes and aromatic compounds, with small amounts of N, O and S compounds. It formed from marine organisms. It is refined into fuels (petrol, diesel, kerosene) and is also a vital **chemical feedstock** for polymers, pharmaceuticals, dyes and solvents.
- **Natural gas** — mainly **methane**, with some ethane, nitrogen and hydrogen sulfide impurities. It is often found above oil or with coal, trapped under impermeable rock.

4.3 Advantages and disadvantages

	Coal	Crude oil (petroleum)	Natural gas
Advantages	cheap and the most plentiful; widely distributed; longest remaining lifespan; can be converted into synthetic liquid and gaseous fuels	liquid, so easily transported by pipeline or tanker; high energy density, ideal for vehicles; volatile and burns easily; sulfur impurities can be removed in refining	highest specific energy of the three; cleanest: lowest %C, so least CO_2, CO and soot per kJ; easily piped; sulfur removed easily, so no acid rain
Disadvantages	most CO_2 per kJ; SO_2 from sulfur impurities causes acid rain; particulates; bulky solid, hard to transport; mining is dangerous and damages landscapes	limited reserves, unevenly distributed (political and economic issues); CO_2, CO and NO_x from engines; photochemical smog; spills pollute oceans	limited reserves; still produces CO_2; methane itself is a potent greenhouse gas if it leaks; risk of explosion; pipelines need very large investment

The key idea in one line

All fossil fuels are **non-renewable** and all produce the greenhouse gas **carbon dioxide**. The **higher the carbon content**, the **more CO_2 per kJ** and the **greater the tendency for incomplete combustion**: natural gas is the cleanest, coal the dirtiest.

5 Reactivity 1.3.3 — Comparing fuels: energy and CO_2 SL + HL

5.1 Specific energy and energy density

Definitions to learn

Specific energy is the energy released per unit **mass** of fuel (kJ g^{-1} or MJ kg^{-1}) = $|\Delta H_c|$ / M.

Energy density is the energy released per unit **volume** of fuel (kJ cm^{-3} or MJ dm^{-3}) = specific energy × density.

Fuel	ΔH_c / kJ mol^{-1}	specific energy / kJ g^{-1}	energy density / kJ cm^{-3}	Best suited to
hydrogen, $H_2(g)$	−286	**141.6**	0.0115	rockets, where mass matters most
methane, $CH_4(g)$	−891	55.5	0.0359	pipelines, domestic heating
octane, $C_8H_{18}(l)$	−5470	47.9	**33.7**	cars, where tank volume matters
ethanol, $C_2H_5OH(l)$	−1367	29.7	23.4	biofuel blended with petrol
methanol, $CH_3OH(l)$	−726	22.7	18.0	fuel cells, racing fuel

Gas energy densities are at 298 K and 100 kPa (molar volume 24.8 dm^3); liquid densities: octane 0.703, ethanol 0.789, methanol 0.792 g cm^{-3}.

Gases have very high specific energies but very low energy densities, because a mole of gas occupies a large volume. Liquids are far better per unit volume, which is why petrol and diesel dominate transport. Alcohols are less energetic than alkanes per gram because they are **already partly oxidized**: the oxygen atom adds mass but releases no energy.

Worked example 8 — specific energy and energy density

Q. (a) Calculate the specific energy of ethanol. (b) Calculate the volume of octane, and the volume of hydrogen gas at 298 K and 100 kPa, needed to release 1000 kJ.

(a) 1367 kJ mol^{-1} / 46.08 g mol^{-1} = **29.7 kJ g^{-1}**

(b) octane: 1000 / 33.7 = **29.7 cm^3**

hydrogen: n = 1000 / 286 = 3.50 mol; V = 3.50 × 24.8 = **86.7 dm^3**, almost 3000 times the volume of octane

This is the main practical problem with hydrogen as a vehicle fuel: it must be compressed to very high pressures or liquefied at about 20 K.

5.2 How much CO_2 does a fuel produce?

The syllabus asks you to **evaluate** the CO_2 each fuel adds. The useful question is not simply "how much CO_2 per gram of fuel?" but "how much CO_2 for the energy we get?". A useful measure is the mass of CO_2 per 1000 kJ released (or, equivalently, the energy released per gram of CO_2).

Worked example 9 — CO_2 from pentane

Q. Calculate the mass of CO_2 produced (a) when 1.00 g of pentane burns completely and (b) per 1000 kJ of energy released. $\Delta H_c(C_5H_{12}) = -3509$ kJ mol^{-1}.

$C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$ · $M(C_5H_{12}) = 72.17$ g mol^{-1}

(a) $n(CO_2) = 5 \times 1.00/72.17 = 0.0693$ mol; $m = 0.0693 \times 44.01 =$ **3.05 g**

(b) 1000 kJ needs $1000/3509 = 0.285$ mol pentane → $5 \times 0.285 = 1.425$ mol $CO_2 =$ **62.7 g**

Alkane	CH_4	C_2H_6	C_3H_8	C_4H_{10}	C_5H_{12}	C_6H_{14}	C_7H_{16}	C_8H_{18}
% carbon by mass	74.8	79.9	81.7	82.6	83.2	83.6	83.9	84.1
g CO_2 per g of fuel	2.74	2.93	2.99	3.03	3.05	3.06	3.07	3.08
specific energy / kJ g^{-1}	55.5	51.9	50.3	49.5	48.6	48.3	48.1	47.9
g CO_2 per 1000 kJ	**49.4**	56.4	59.5	61.2	62.7	63.4	64.0	64.4

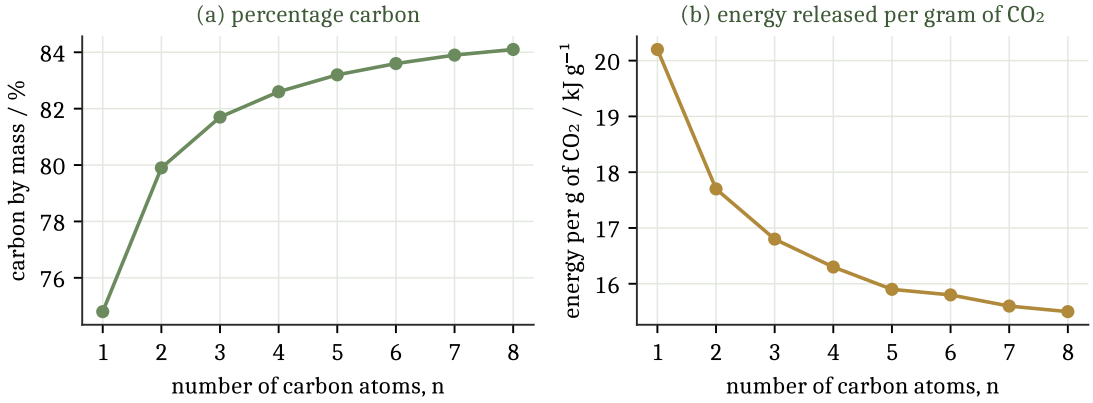


Figure 2 — Calculated from data booklet ΔH_c values. As the chain lengthens, %C rises and the energy obtained per gram of CO_2 falls: methane gives the most energy for the least CO_2.

Two trends follow from the table. A **higher percentage of hydrogen** gives a higher specific energy, because forming O–H bonds releases a lot of energy for very little mass; and a **higher percentage of carbon** gives more CO_2 per kJ. Coal, which is close to pure carbon, is the worst: C(s) gives 44.01 g of CO_2 for 394 kJ, i.e. 112 g per 1000 kJ.

Worked example 10 — a coal-fired power station ★

Q. A 600 MW power station (600×10^3 kJ s^{-1} of electrical output) burns coal of specific energy 30.0 kJ g^{-1} with an overall efficiency of 40 %. The coal can be represented as CH. Calculate the mass of (a) coal burned and (b) CO_2 released each second.

heat needed per second = $600 \times 10^3 / 0.40 = 1.50 \times 10^6$ kJ

(a) mass of coal = $1.50 \times 10^6 / 30.0 = 5.00 \times 10^4$ g = **50.0 kg s^{-1}**

(b) n(CH) = $5.00 \times 10^4 / 13.02 = 3840$ mol → 3840 mol $CO_2 \times 44.01 = 1.69 \times 10^5$ g = **169 kg s^{-1}** (over 14 000 tonnes a day)

Examiner traps in comparing fuels

- Confusing specific energy (per mass) with energy density (per volume).
- Comparing ΔH_c per mole when the question asks which fuel is "better": always convert to per gram, per cm^3 or per gram of CO_2, whichever the context needs.
- Forgetting the efficiency: the heat that must be released = useful output ÷ efficiency.

6 Reactivity 1.3.3 — CO_2 and the greenhouse effect SL + HL

6.1 Rising carbon dioxide levels

Atmospheric CO_2 has been measured continuously at the Mauna Loa Observatory in Hawaii since 1958. It has risen steadily from about 315 ppm to over 420 ppm, mainly because of the burning of fossil fuels, with deforestation adding to the increase. Superimposed on the rise is a regular **seasonal fluctuation**: CO_2 falls each northern spring and summer as plants grow and photosynthesize, and rises each autumn and winter as plants decay and photosynthesis slows. Over the same period, average global temperatures have also risen.

6.2 How the greenhouse effect works

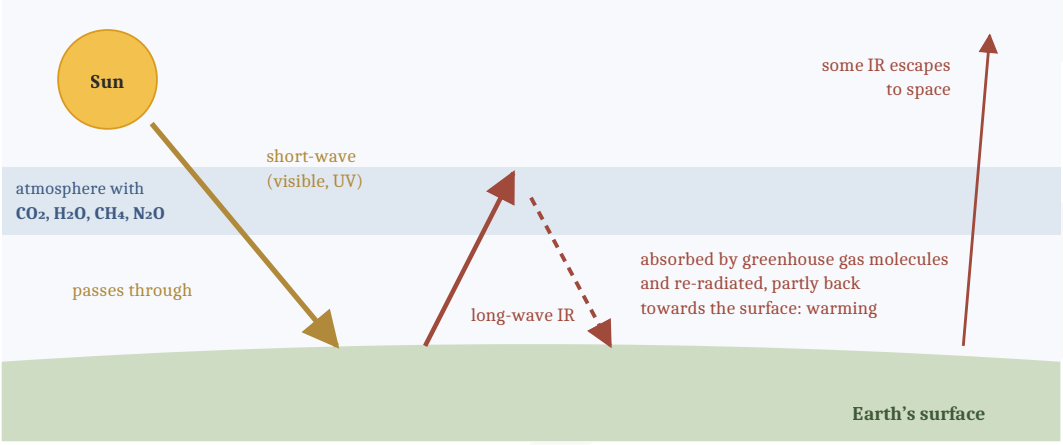


Figure 3 — The greenhouse effect. Short-wavelength radiation from the Sun passes through the atmosphere; longer-wavelength infrared radiation from the warm Earth is absorbed by greenhouse gases and partly re-radiated back towards the surface.

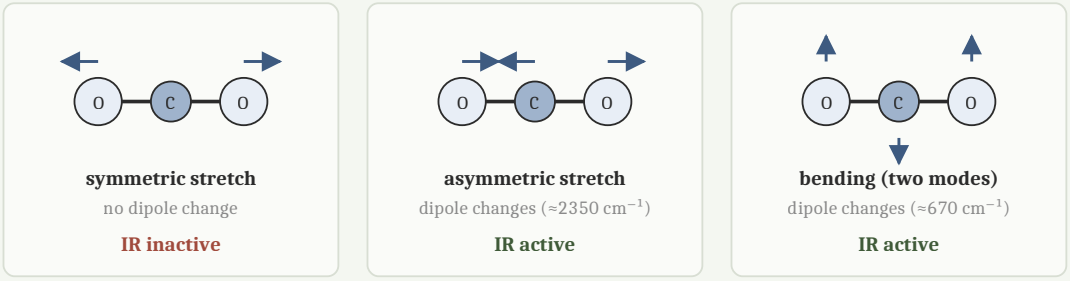
1. Short-wavelength (visible and near-UV) radiation from the Sun passes through the atmosphere and warms the Earth's surface.
2. The warm surface radiates energy back as longer-wavelength **infrared** radiation.
3. **Greenhouse gases** (CO_2, H_2O, CH_4, N_2O) absorb some of this IR radiation. Their bonds vibrate more energetically.
4. The molecules re-radiate the energy in all directions, including back towards the surface, so the lower atmosphere warms.

The natural greenhouse effect keeps the Earth about 33 °C warmer than it would otherwise be and is essential for life. The concern is the **enhanced** greenhouse effect: extra CO_2 and other gases from human activities trap more heat, causing global warming. Consequences include changes in rainfall and crop yields, desertification, loss of habitats, more extreme weather, and rising sea levels from the thermal expansion of seawater and melting ice.

6.3 Why CO_2 absorbs infrared radiation

HL link: Structure 3.2

A molecule can only absorb IR radiation if the vibration causes a **change in its dipole moment**. Linear O=C=O has no permanent dipole, but two of its three types of vibration create a temporary one:



N_2 and O_2, which make up 99 % of the air, are not greenhouse gases: their vibrations (simple stretching of two identical atoms) never change the dipole moment, so they cannot absorb IR radiation.

6.4 Correlation and cause, and carbon footprints

The rise in CO_2 and the rise in temperature are **correlated**, but a correlation alone does not prove a cause: both could depend on a third variable. The case for a causal link rests on a plausible mechanism (IR absorption, above) supported by computer models, because CO_2 levels cannot be changed in a controlled experiment on the whole planet. A **carbon footprint** is a measure of the total greenhouse gases produced by an individual, product or activity, expressed as a mass of CO_2 equivalent. There is debate over exactly what it should include, for example whether methane and N_2O should be counted, and how.

Model answer: "Explain how increasing concentrations of carbon dioxide in the atmosphere lead to global warming." (3 marks)

Short-wavelength radiation from the Sun passes through the atmosphere and is absorbed by the Earth's surface, which re-emits longer-wavelength infrared radiation **[1]**. CO_2 molecules absorb this IR radiation because their bending and asymmetric stretching vibrations change the dipole moment **[1]**. The energy is re-radiated in all directions, including back towards the Earth, so more CO_2 traps more heat and the average temperature rises **[1]**.

Examiner traps on the greenhouse effect

- Mixing up global warming with **ozone depletion**. They are separate problems: holes in the ozone layer do not let "more heat in".
- Saying CO_2 absorbs radiation from the Sun. It absorbs the **infrared** radiation emitted by the Earth.
- Forgetting "re-radiated back towards the Earth": absorption alone does not explain warming.
- Calling the greenhouse effect itself harmful. It is the **enhanced** effect that causes concern.

7 Reactivity 1.3.4 — Biofuels SL + HL

7.1 Photosynthesis: the source of all biomass

Green plants use the pigment **chlorophyll** to absorb sunlight and drive a sequence of redox reactions that convert carbon dioxide and water into glucose and oxygen:

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

light energy is converted into chemical energy · the reverse of the combustion (and respiration) of glucose

Photosynthesis is not very efficient: only a small fraction (typically 1–2 %) of the sunlight reaching a plant is stored as chemical energy, because much of the light is reflected, transmitted, or of the wrong wavelength, and plants use energy for their own respiration.

7.2 Renewable and non-renewable sources

Definitions to learn

A **renewable** energy source is replaced naturally on a human timescale, as fast as it is used, so it will not run out (biofuels, solar, wind, hydroelectric).

A **non-renewable** energy source is finite: it is used much faster than it can be replaced (fossil fuels, which take millions of years to form; uranium).

Biofuels are fuels produced from **biomass** (recently living matter): carbon is fixed from the atmosphere by photosynthesis over a **short period of time**, months or years rather than millions of years.

7.3 Types of biofuel

Biofuel	How it is made	Notes
wood and other biomass	burned directly; wood is mainly cellulose, a polymer of glucose	lower specific energy than coal (about 15–18 kJ g^{-1}) because of its high oxygen content; burning in an enclosed space produces CO, soot and unburnt hydrocarbons
ethanol (bioethanol)	**fermentation** of sugars from crops such as sugar cane or maize by **yeast**, which provides the enzymes: $C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$ about 35–37 °C, **absence of oxygen** (anaerobic), aqueous; then fractional distillation	used neat or blended with petrol ("gasohol", e.g. 10 % ethanol). Lower CO and NO_x emissions; reduces dependence on oil. But it absorbs water (hydrogen bonding), which can cause separation from the hydrocarbon fuel and corrosion
biogas (mainly methane)	bacterial (anaerobic) decomposition of plant waste, manure or sewage: $C_6H_{12}O_6 \rightarrow 3CO_2 + 3CH_4$	a mixture of roughly 50–70 % CH_4 with CO_2; useful in rural areas and for waste disposal
biodiesel	vegetable oils converted into esters (Structure 3.2 link)	can be used in diesel engines; uses land that could grow food

7.4 Are biofuels carbon neutral?

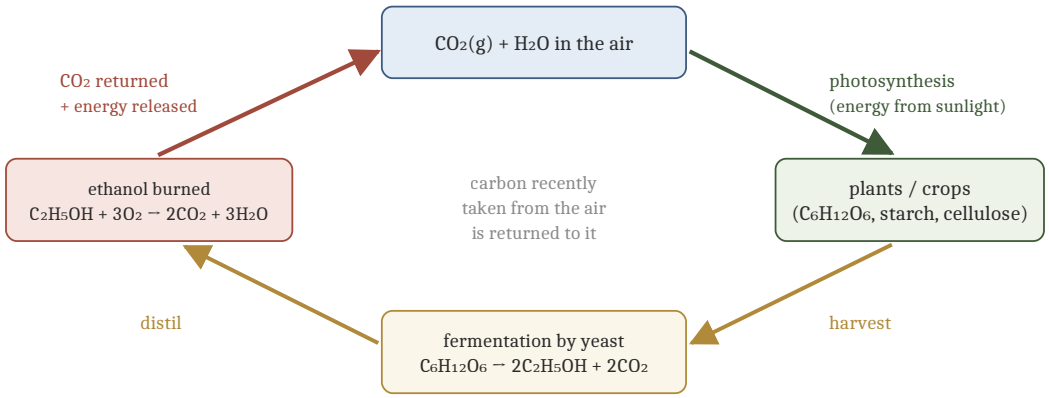


Figure 4 — The carbon cycle of bioethanol. The CO_2 released when the ethanol burns was absorbed from the air by the crop only months earlier.

In principle a biofuel is **carbon neutral**: the CO_2 released on burning equals the CO_2 absorbed during photosynthesis. In practice it is not, because energy from fossil fuels is used to make fertilizers, run farm machinery, transport the crop, and distil the ethanol. Biofuels have a much smaller net carbon footprint than fossil fuels, but not zero.

Advantages of biofuels	Disadvantages of biofuels
renewable and sustainable if crops are replanted	use land that could grow food, raising food prices; may lead to deforestation
lower net CO_2 emissions than fossil fuels	large volumes needed; harvesting and transport are costly
less polluting (fewer sulfur impurities, less CO, NO_x)	remove nutrients from soil; need fertilizers and lots of water
reduce dependence on imported oil	lower specific energy than fossil fuels
can use waste materials (biogas)	not truly carbon neutral

Worked example 11 — fermentation

Q. (a) Calculate the maximum mass of ethanol that can be made from 1.00 kg of glucose. (b) Use ΔH_c values (glucose −2803, ethanol −1367 kJ mol^{-1}) to find the percentage of the energy in the glucose that is retained in the ethanol.

(a) n(glucose) = 1000/180.18 = 5.55 mol → 11.1 mol C_2H_5OH × 46.08 = **511 g**

(b) energy in ethanol from 1 mol glucose = 2 × 1367 = 2734 kJ; 2734 / 2803 × 100 = **97.5 %**

Fermentation itself wastes little energy (ΔH = −69 kJ mol^{-1}, from Booklet 2); the real energy costs of bioethanol are in farming and distillation.

Model answer: "Explain why biofuels are described as renewable and discuss whether they are carbon neutral." (3 marks)

They come from crops that can be regrown within a short time, fixing carbon from the atmosphere by photosynthesis, so they are replaced as fast as they are used **[1]**. The CO_2 released on combustion was recently absorbed by the plants, so in principle there is no net increase **[1]**. However, fossil fuels are used in growing, fertilizing, transporting and processing the crops, so there is still a net release of CO_2 **[1]**.

8 Reactivity 1.3.5 — Fuel cells SL + HL

8.1 Electricity without combustion

The combustion of hydrogen, $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$, $\Delta H = -286$ kJ mol^{-1}, is a redox reaction in which electrons are transferred from hydrogen to oxygen. If the two reactants are kept apart and supplied to separate electrodes, the electrons can be made to flow through an external circuit instead of the energy being released as heat. This is a **fuel cell**: a device that converts chemical energy from a fuel **directly into electrical energy**.

The essentials

- The **fuel** (H_2, CH_3OH) is supplied continuously to the **anode** (negative electrode), where it is **oxidized**: electrons are released.
- **Oxygen** (or air) is supplied continuously to the **cathode** (positive electrode), where it is **reduced**: electrons are gained.
- Electrons flow through the external circuit from anode to cathode; ions move through the **electrolyte** to complete the circuit.
- The electrodes are porous, often made of carbon, and contain a catalyst: platinum, or transition metals such as nickel.
- The cell keeps working as long as fuel and oxygen are supplied, because the products are removed continuously.

8.2 The hydrogen–oxygen fuel cell

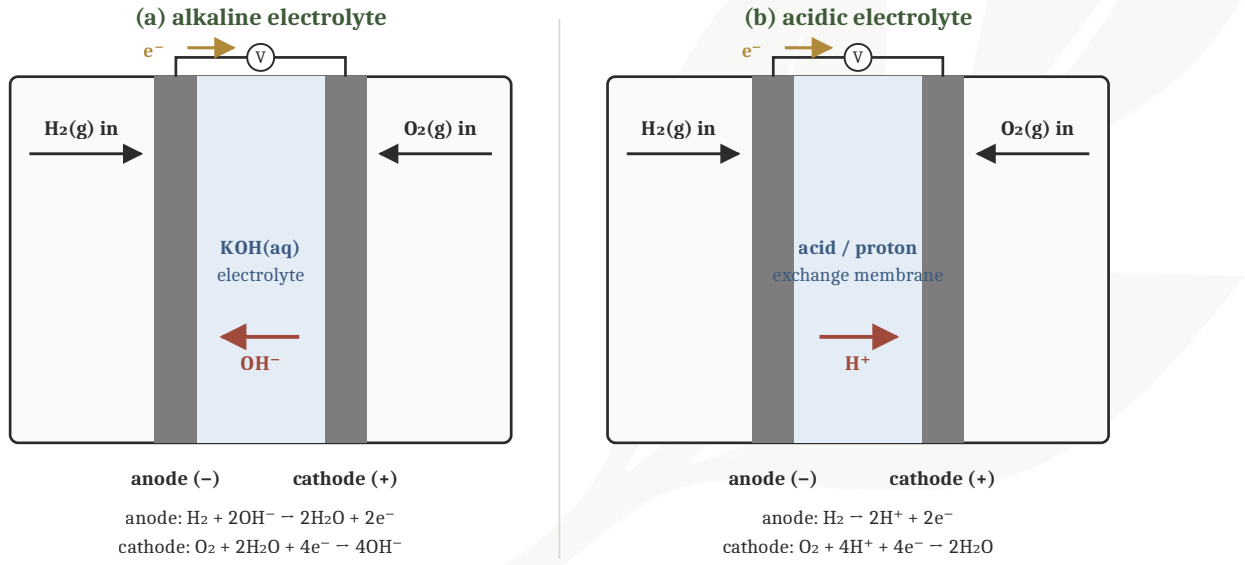


Figure 5 — Hydrogen–oxygen fuel cells. (a) In an alkaline electrolyte, OH^- ions move to the anode and water forms there. (b) In an acidic electrolyte, H^+ ions move to the cathode and water forms there. The overall reaction is the same in both. (Details of the proton exchange membrane are not assessed.)

Electrolyte	Anode (−): oxidation	Cathode (+): reduction
alkaline (KOH(aq))	$H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$	$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$
acidic	$H_2(g) \rightarrow 2H^+(aq) + 2e^-$	$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$
overall (both)	$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ (anode equation × 2, then add; electrons cancel)	

8.3 Deducing half-equations

Method — balancing half-equations in acid, then converting to alkali

1. Balance the atoms other than O and H.
2. Balance **O** by adding H_2O.
3. Balance **H** by adding H^+.
4. Balance **charge** by adding electrons (on the left for reduction, on the right for oxidation).
5. For an **alkaline** electrolyte: add as many OH^- to both sides as there are H^+; combine $H^+ + OH^-$ into H_2O; cancel water.

Worked example 12 — the methanol fuel cell (acidic electrolyte)

Q. Methanol is oxidized to carbon dioxide at the anode. Deduce both half-equations and the overall equation.

Anode: $CH_3OH \rightarrow CO_2$

O: 1 on the left, 2 on the right → add 1 H_2O on the left: $CH_3OH + H_2O \rightarrow CO_2$

H: 6 on the left → add $6H^+$ on the right; charge: add $6e^-$ on the right

$CH_3OH(aq) + H_2O(l) \rightarrow CO_2(g) + 6H^+(aq) + 6e^-$

Cathode (6 electrons needed): **$1\frac{1}{2}O_2(g) + 6H^+(aq) + 6e^- \rightarrow 3H_2O(l)$**

Overall: $CH_3OH + H_2O + 1\frac{1}{2}O_2 \rightarrow CO_2 + 3H_2O$, i.e. **$CH_3OH(aq) + 1\frac{1}{2}O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$**

The overall reaction is the same as the combustion of methanol, but the energy is released as electricity. The oxidation state of carbon rises from −2 to +4: a change of six, matching the six electrons.

Worked example 13 — converting the hydrogen cathode equation to alkaline conditions

acid: $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$

add $4OH^-$ to both sides: $O_2 + 4H^+ + 4OH^- + 4e^- \rightarrow 2H_2O + 4OH^-$

$4H^+ + 4OH^- = 4H_2O$: $O_2 + 4H_2O + 4e^- \rightarrow 2H_2O + 4OH^-$

cancel $2H_2O$: **$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$**

Worked example 14 — the oxidation-state method (alkaline hydrogen fuel cell)

An alternative to the acid-then-alkali method: use oxidation states to find the number of electrons first, then balance charge with the ion that is present in the electrolyte.

Step	Anode: H_2 is oxidized	Cathode: O_2 is reduced
1 · skeleton and oxidation states	$H_2 \rightarrow H_2O$ H: 0 → +1	$O_2 \rightarrow H_2O$ O: 0 → −2
2 · electrons	2 H atoms each lose 1 e^-: $H_2 \rightarrow H_2O + 2e^-$	2 O atoms each gain 2 e^-: $O_2 + 4e^- \rightarrow 2H_2O$
3 · balance charge with OH^-	add $2OH^-$ on the left	add $4OH^-$ on the right
4 · balance H and O with H_2O	**$H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$**	$O_2 + 4H_2O + 4e^- \rightarrow 2H_2O + 4OH^-$, simplified: **$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$**

In an acidic electrolyte, balance the charge with H^+ instead. For methanol, carbon changes from −2 to +4, which gives the 6 electrons in worked example 12 straight away.

8.4 Fuel cells compared with combustion and with primary cells

Advantages of fuel cells	Disadvantages of fuel cells
more efficient (about 40–60 %) than burning fuel in an engine or power station (about 20–40 %), because chemical energy is converted directly to electrical energy without the losses of a heat engine the hydrogen cell's only product is water: no CO_2, CO, particulates or NO_x at the point of use quiet, few moving parts; run continuously while fuel is supplied	hydrogen is difficult to store and transport (very low energy density; flammable/explosive) most hydrogen is currently made from fossil fuels (e.g. methane and steam), producing CO_2; it is only "green" if made by electrolysis using renewable electricity platinum catalysts are expensive; methanol cells still release CO_2, and methanol is toxic

	Fuel cell	Primary (voltaic) cell, e.g. a zinc–carbon battery
reactants	supplied **continuously** from outside	sealed inside the cell; a fixed, finite amount
lifetime	works for as long as fuel and oxygen are supplied	"runs flat" when the reactants are used up; not rechargeable
products	removed continuously (e.g. water)	stay inside the cell
common to both	a spontaneous redox reaction; oxidation at the anode (−), reduction at the cathode (+); electrons flow through the external circuit; chemical energy → electrical energy	

Beyond the syllabus: microbial fuel cells use bacteria to oxidize organic matter, including waste, and harvest some of the energy as electricity. They are being researched as a cheap, renewable alternative to hydrogen cells, but their details are not assessed.

Model answer: "State the half-equations at each electrode in a hydrogen–oxygen fuel cell with an alkaline electrolyte and identify the electrode at which oxidation occurs." (3 marks)

Anode: $H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$ **[1]**. Cathode: $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$ **[1]**. Oxidation (loss of electrons) occurs at the anode, the negative electrode **[1]**.

Examiner traps in 1.3.5

- Putting H^+ in an alkaline half-equation, or OH^- in an acidic one.
- Getting the electrons on the wrong side: oxidation releases electrons (right), reduction gains them (left).
- Saying a fuel cell “burns” hydrogen. There is no combustion; the energy is released as electricity.
- Saying hydrogen fuel cells produce “no pollution at all”: making and storing the hydrogen may release CO_2.

9 Key points to memorise

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

A Equations

- $C_xH_y + (x + y/4)O_2 \rightarrow xCO_2 + y/2\ H_2O$ · to CO: $(x/2 + y/4)O_2 \rightarrow xCO + y/2\ H_2O$ · remember the O in alcohols.
- Photosynthesis: $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$ · Fermentation: $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$ (yeast, ~37 °C, no O_2).
- H_2 fuel cell, acid: $H_2 \rightarrow 2H^+ + 2e^-$; $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$. Alkali: $H_2 + 2OH^- \rightarrow 2H_2O + 2e^-$; $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$.
- Methanol fuel cell, acid: $CH_3OH + H_2O \rightarrow CO_2 + 6H^+ + 6e^-$; $1½O_2 + 6H^+ + 6e^- \rightarrow 3H_2O$.
- Catalytic converter: $2CO + O_2 \rightarrow 2CO_2$; $2CO + 2NO \rightarrow 2CO_2 + N_2$.

B Definitions and calculations

- **Specific energy** = energy per unit mass = $|\Delta H_c|/M$. **Energy density** = energy per unit volume.
- **Renewable:** replaced as fast as it is used. **Biofuel:** from biomass, carbon fixed by photosynthesis over a short time.
- **Fuel cell:** converts chemical energy of a fuel directly into electrical energy; fuel oxidized at the anode (−), O_2 reduced at the cathode (+).
- CO_2 per kJ: find mol of fuel for the energy, multiply by number of C atoms, × 44.01.

C Explanations you will be asked for

- Incomplete combustion: limited O_2 or high %C → CO and C; less energy released; CO binds to haemoglobin; soot harms lungs.
- Higher %C → sootier flame and more CO_2 per kJ: natural gas cleanest, coal dirtiest. Higher %H → higher specific energy.
- Greenhouse effect: IR from the Earth absorbed by CO_2 (vibrations change dipole) and re-radiated back. N_2, O_2 are not greenhouse gases.
- Fuels need a high E_a for safe storage (kinetically stable). Alcohols have lower specific energy: already partly oxidized.
- Biofuels are renewable but not truly carbon neutral; they compete with food for land.

10 Check yourself

Use the data in this booklet where needed: ΔH_c / kJ mol^{-1}: CH_4 −891, C_3H_8 −2219, C_8H_{18} −5470, CH_3OH −726, C_2H_5OH −1367, H_2 −286. Marks are shown in brackets; allow about 1.2 minutes per mark. Full answers are in section 11.

Part A — Multiple choice SL + HL

1. Which oxide, formed by combustion of the element, dissolves in water to give an acidic solution?

A Na_2O **B** MgO **C** CaO **D** SO_2

2. What is the coefficient of O_2 when the equation for the complete combustion of one mole of propan-1-ol is balanced?

A 3 **B** 4 **C** 4½ **D** 5

3. One mole of which compounds requires exactly 3 mol of O_2 for complete combustion?

I. C_2H_4 II. C_2H_5OH III. CH_3COOH

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

4. Which is the correct equation for the incomplete combustion of propane to carbon monoxide?

A $C_3H_8 + 2O_2 \rightarrow 3CO + 4H_2O$ **B** $2C_3H_8 + 7O_2 \rightarrow 6CO + 8H_2O$
C $C_3H_8 + 5O_2 \rightarrow 3CO + 4H_2O$ **D** $2C_3H_8 + 7O_2 \rightarrow 6CO + 8H_2$

5. Why is carbon monoxide toxic?

A It dissolves in rain to form an acid.
B It binds to haemoglobin more strongly than oxygen does.
C It absorbs infrared radiation.
D It is a solid particulate that damages the lungs.

6. Which compound burns with the sootiest flame?

A methane **B** ethanol **C** hexane **D** benzene

7. Why does a mixture of methane and air not ignite at room temperature?

A The combustion of methane is endothermic.
B The reaction has a high activation energy.
C Methane is more stable than CO_2 and H_2O.
D Oxygen is not an oxidizing agent at room temperature.

8. Which fuel has the highest specific energy?

A hydrogen **B** methane **C** octane **D** ethanol

9. Which fuel releases the **smallest** mass of CO_2 per kJ of energy released?

A coal **B** octane **C** methane **D** ethanol

10. Why are liquid fuels such as petrol preferred to hydrogen in cars?

A They have a higher specific energy.
B They have a much higher energy density.
C They produce no carbon dioxide.
D They have a lower carbon content.

11. Which gas is **not** a greenhouse gas?

A N_2 **B** CH_4 **C** H_2O **D** N_2O

12. Which describes how carbon dioxide contributes to global warming?

A It absorbs ultraviolet radiation from the Sun.
B It absorbs infrared radiation emitted by the Earth and re-radiates some back towards the surface.
C It destroys the ozone layer.
D It reflects visible light back to the Earth.

13. Which are the conditions for making ethanol by fermentation?

A yeast, about 37 °C, absence of oxygen
B yeast, about 37 °C, excess oxygen
C phosphoric acid catalyst, 300 °C, steam
D platinum catalyst, room temperature, oxygen

14. Which statement explains why biofuels are renewable?

A They release no carbon dioxide when burned.
B The crops they are made from can be regrown in a short time.
C They have a higher specific energy than fossil fuels.
D They are formed over millions of years.

15. What is the anode half-equation in a hydrogen–oxygen fuel cell with an alkaline electrolyte?

A $H_2 \rightarrow 2H^+ + 2e^-$ **B** $H_2 + 2OH^- \rightarrow 2H_2O + 2e^-$
C $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$ **D** $2H^+ + 2e^- \rightarrow H_2$

16. How many electrons are released per molecule of methanol oxidized at the anode of a methanol fuel cell?

A 2 **B** 4 **C** 6 **D** 8

17. Which is a difference between a fuel cell and a primary cell?

A Only the fuel cell involves a redox reaction.
B Only the fuel cell has its reactants supplied continuously from outside.
C Only the primary cell converts chemical energy into electrical energy.
D In a fuel cell, oxidation occurs at the cathode.

Part B — Structured questions SL + HL

18. Write balanced equations, with state symbols, for: **[7]**

(a) the combustion of sodium; (b) the combustion of white phosphorus to phosphorus(V) oxide;

(c) the complete combustion of butan-1-ol, $C_4H_9OH(l)$; (d) the complete combustion of heptane, $C_7H_{16}(l)$;

(e) the incomplete combustion of ethane to carbon; (f) the incomplete combustion of ethanol to carbon monoxide;

(g) the general equation for the complete combustion of an alkene C_nH_{2n}.

19. 0.601 g of a compound containing C, H and O burned completely to give 1.320 g of CO_2 and 0.721 g of H_2O. Its molar mass is 60.1 g mol^{-1}.

(a) Determine the empirical formula. **[3]**

(b) Deduce the molecular formula and draw two possible structural isomers. **[2]**

20. Car engines can burn octane incompletely.

(a) Write an equation for the incomplete combustion of octane to carbon monoxide and steam. [1]

(b) Calculate the mass of carbon monoxide formed when 1.00 g of octane burns in this way. [2]

(c) Explain why carbon monoxide is a dangerous pollutant. [2]

(d) Nitrogen monoxide also forms in engines. Write an equation for its formation and suggest why it forms. [2]

(e) Write an equation for the removal of both pollutants in a catalytic converter. [1]

21. (HL) Use ΔH_f / kJ mol^{-1}: $CH_4(g)$ −75, $CO(g)$ −111, $CO_2(g)$ −394, $H_2O(l)$ −286.

(a) Calculate ΔH for $CH_4(g) + 1½O_2(g) \rightarrow CO(g) + 2H_2O(l)$. [2]

(b) Calculate the percentage of the energy of complete combustion (−891 kJ mol^{-1}) that is "lost" when CO forms instead of CO_2. [1]

(c) Explain, in terms of oxidation, why less energy is released. [1]

22. The table gives approximate data for five fuels.

Fuel	specific energy / kJ g^{-1}	carbon content by mass / %
coal	32	94
oil	42	83
natural gas	55	75
wood	15	50
hydrogen	142	0

(a) Suggest an explanation for the trend in specific energy from coal to hydrogen. [1]

(b) Calculate the mass of CO_2 produced for 1000 kJ of energy from each fuel and identify the best and worst fossil fuels on this basis. [3]

(c) Wood gives the most CO_2 per kJ, yet is often considered a better choice than coal. Explain why. [1]

23. ★ A 500 MW power station burns coal with a specific energy of 30.0 kJ g^{-1} and a carbon content of 85 % by mass, at an efficiency of 35 %. Calculate the mass of (a) coal burned and (b) CO_2 produced each second. [4]

24. ★ A home needs 1.20×10^6 kJ of heat on a winter day. It could burn coal (formula approximately CH, specific energy 30 kJ g^{-1}, heating efficiency 60 %) or wood (approximately $C_6H_{10}O_5$, specific energy 17 kJ g^{-1}, efficiency 75 %).

(a) Calculate the percentage of carbon by mass in each fuel. [1]

(b) Calculate the mass of each fuel needed and the mass of CO_2 each produces. [3]

(c) Suggest why the CO_2 calculation alone does not give a full picture of the environmental impact of either fuel. [2]

25. (a) Calculate the specific energies of methanol and octane and explain why methanol's is lower. [3]

(b) Octane has a density of 0.703 g cm^{-3}. Calculate the volume of octane, and of hydrogen gas at 298 K and 100 kPa (molar volume 24.8 dm^3 mol^{-1}), that would each release 1000 kJ. [2]

(c) Use your answers to explain why hydrogen is used as a rocket fuel but rarely as a car fuel. [2]

26. (a) Describe how carbon dioxide in the atmosphere causes the greenhouse effect. [3]

(b) (HL) Explain why CO_2 absorbs infrared radiation but N_2 does not. [2]

(c) The CO_2 concentration at Mauna Loa rises and falls each year. Explain this fluctuation. [1]

(d) State two possible consequences of global warming. [2]

(e) Explain why a correlation between CO_2 levels and temperature does not on its own prove that one causes the other. [1]

27. Bioethanol is made from sugar cane.

(a) Write the equation for photosynthesis and state the name of the pigment that absorbs light. **[2]**

(b) Write the equation for fermentation and state the conditions. **[3]**

(c) Calculate the maximum mass of ethanol obtainable from 500 g of glucose. **[2]**

(d) Explain why bioethanol is described as renewable, and discuss whether it is carbon neutral. **[3]**

(e) State two disadvantages of growing crops for biofuels. **[2]**

28. ★ Biogas is formed by the anaerobic decomposition of biomass: $C_6H_{12}O_6 \rightarrow 3CO_2 + 3CH_4$.

(a) Calculate the energy released by burning the methane produced from one mole of glucose. **[1]**

(b) Compare this with ΔH_c of glucose (-2803 kJ mol^{-1}) and comment. **[2]**

29. Fuel cells.

(a) Write the half-equations for a hydrogen–oxygen fuel cell with an acidic electrolyte, identifying the anode and cathode.

(b) Write the half-equations for a methanol fuel cell with an acidic electrolyte and the overall equation. **[3][2]**

(c) State two advantages of a hydrogen fuel cell over a petrol engine and one disadvantage. **[3]**

(d) State one similarity and one difference between a fuel cell and a primary cell. **[2]**

30. ★ Ethanol can also be used in a fuel cell with an acidic electrolyte, where it is oxidized to carbon dioxide.

(a) Deduce the anode half-equation. **[2]**

(b) Deduce the cathode half-equation that matches it and the overall equation. **[2]**

(c) State the change in the average oxidation state of carbon and show that it is consistent with the number of electrons. **[1]**

31. ★ A student suggests that the more reduced the carbon in a fuel, the higher its specific energy.

(a) Deduce the average oxidation state of carbon in methane, methanol and glucose, $C_6H_{12}O_6$. **[1]**

(b) Calculate the specific energies of methane, methanol and glucose (ΔH_c of glucose = -2803 kJ mol^{-1}). **[2]**

(c) Octane (oxidation state -2.25) has a specific energy of 47.9 kJ g^{-1} and ethanol (-2) has 29.7 kJ g^{-1}. Evaluate the student's hypothesis using all five fuels. **[2]**

Part A: 17 marks · Part B: 90 marks · Total: 107 marks (about 2 hours)

11 Answers

Part A

1. D. Non-metal oxides are acidic: $SO_2 + H_2O \rightarrow H_2SO_3$. The metal oxides are basic.

2. C. $C_3H_7OH + 4\frac{1}{2}O_2 \rightarrow 3CO_2 + 4H_2O$ (10 O on the right, minus 1 in the alcohol).

3. A. C_2H_4: 2 + 1 = 3; C_2H_5OH: 3; CH_3COOH needs only 2 (it already contains 2 O).

4. B. 6 + 8 = 14 O atoms on the right = $7O_2$. A is unbalanced in O; D makes H_2.

5. B.

6. D. Benzene has the highest %C (92.2 %).

7. B. The fuel is kinetically stable.

8. A. 141.6 kJ g^{-1}.

9. C. Methane has the lowest %C: 49.4 g CO_2 per 1000 kJ, against about 64 for octane and ethanol and over 100 for coal.

10. B. Hydrogen has the higher *specific* energy, but as a gas a very low energy density.

11. A. N_2 vibrations do not change the dipole moment.

12. B.

13. A. C describes the industrial hydration of ethene.

14. B. They do release CO_2 (A is wrong); D describes fossil fuels.

15. B. A is the acidic anode; C is the alkaline cathode.

16. C. $CH_3OH + H_2O \rightarrow CO_2 + 6H^+ + 6e^-$; carbon goes from −2 to +4.

17. B. Both involve redox and convert chemical to electrical energy; oxidation is always at the anode.

Part B

18. One mark each (formulas, balancing and states):
(a) $4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$ · (b) $P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$
(c) $C_4H_9OH(l) + 6O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$ · (d) $C_7H_{16}(l) + 11O_2(g) \rightarrow 7CO_2(g) + 8H_2O(l)$
(e) $C_2H_6(g) + 1\frac{1}{2}O_2(g) \rightarrow 2C(s) + 3H_2O(l)$ (or $2C_2H_6 + 3O_2 \rightarrow 4C + 6H_2O$) · (f) $C_2H_5OH(l) + 2O_2(g) \rightarrow 2CO(g) + 3H_2O(l)$
(g) $C_nH_{2n} + {}^{3n}/_2O_2 \rightarrow nCO_2 + nH_2O$

19. (a) n(C) = 1.320/44.01 = 0.0300 mol (0.360 g); n(H) = 2 × 0.721/18.02 = 0.0800 mol (0.0808 g) **[1]**; mass O = 0.601 − 0.360 − 0.081 = 0.160 g → 0.0100 mol **[1]**; C : H : O = 3 : 8 : 1 → **C_3H_8O [1]**.
(b) M(C_3H_8O) = 60.1, so the molecular formula is **C_3H_8O [1]**; any two of propan-1-ol $CH_3CH_2CH_2OH$, propan-2-ol $CH_3CH(OH)CH_3$, methoxyethane $CH_3OCH_2CH_3$ **[1]**.

20. (a) $C_8H_{18}(l) + 8\frac{1}{2}O_2(g) \rightarrow 8CO(g) + 9H_2O(g)$ **[1]**.
(b) n(C_8H_{18}) = 1.00/114.26 = 8.75×10^{-3} mol; n(CO) = 8 × that = 0.0700 mol **[1]**; m = 0.0700 × 28.01 = **1.96 g [1]**.
(c) It binds to haemoglobin more strongly than oxygen **[1]**, reducing the oxygen carried by the blood; it is colourless and odourless so it is not detected **[1]**.
(d) $N_2(g) + O_2(g) \rightarrow 2NO(g)$ **[1]**; the very high temperature in the engine supplies the large activation energy (the reaction is endothermic) **[1]**.
(e) $2CO(g) + 2NO(g) \rightarrow 2CO_2(g) + N_2(g)$ **[1]**.

21. (a) [−111 + 2(−286)] − (−75) **[1]** = **−608 kJ mol^{-1} [1]**.
(b) (891 − 608)/891 × 100 = **31.8 % [1]**.
(c) Carbon is only oxidized to the +2 state in CO, not to +4 in CO_2; the energy released when CO is oxidized to CO_2 (−283 kJ) is not obtained **[1]**.

22. (a) Specific energy increases as the percentage of hydrogen increases (carbon content falls); forming O–H bonds releases much energy for little mass (wood is lower because it contains oxygen, so it is partly oxidized) **[1]**.
(b) mass of fuel for 1000 kJ = 1000/specific energy; mass of CO_2 = that × %C × 44.01/12.01: coal **108 g**, oil **72 g**, natural gas **50 g**, wood **122 g**, hydrogen **0 g [2]**; best fossil fuel natural gas, worst coal **[1]**.
(c) Wood is renewable: the CO_2 it releases was recently absorbed by photosynthesis and new trees reabsorb it, so its net contribution is much smaller **[1]**.

23. heat needed = 500 × 10^3/0.35 = 1.43 × 10^6 kJ s^{-1} **[1]**; (a) coal = 1.43 × 10^6/30.0 = 4.76 × 10^4 g = **47.6 kg s^{-1} [1]**; (b) carbon = 0.85 × 4.76 × 10^4 = 4.05 × 10^4 g = 3370 mol **[1]**; CO_2 = 3370 × 44.01 = 1.48 × 10^5 g = **148 kg s^{-1} [1]**.

24. (a) coal CH: 12.01/13.02 = **92.2 %**; wood $C_6H_{10}O_5$: 72.06/162.16 = **44.4 % [1]**.
(b) coal: 1.20 × 10^6/(0.60 × 30) = **66.7 kg**; wood: 1.20 × 10^6/(0.75 × 17) = **94.1 kg [1]**; CO_2 from coal = 66 700/13.02 × 44.01 = **225 kg [1]**; from wood = 94 100/162.16 × 6 × 44.01 = **153 kg [1]**.
(c) Any two: wood is renewable and regrowth reabsorbs CO_2; other greenhouse gases and pollutants (CO, soot, SO_2 from coal, NO_x) are not included; energy used in mining, felling, drying and transport is not included; deforestation reduces CO_2 uptake **[2]**.

25. (a) methanol: 726/32.05 = **22.7 kJ g^{-1}**; octane: 5470/114.26 = **47.9 kJ g^{-1} [2]**; methanol is already partly oxidized: its O atom adds mass without releasing energy (and it has a lower %H) **[1]**.
(b) octane: 47.9 × 0.703 = 33.7 kJ cm^{-3} → **29.7 cm^3**; hydrogen: 1000/286 = 3.50 mol × 24.8 = **86.7 dm^3 [2]**.
(c) In a rocket the mass is what matters, and hydrogen has by far the highest specific energy **[1]**; a car needs a compact fuel tank, and hydrogen's very low energy density would need a huge (or high-pressure) tank **[1]**.

26. (a) short-wavelength radiation from the Sun passes through the atmosphere and warms the Earth **[1]**; the Earth emits longer-wavelength IR radiation, which is absorbed by CO_2 **[1]**; the energy is re-radiated in all directions, partly back to the surface, warming it **[1]**.
(b) Absorption of IR requires a vibration that changes the dipole moment **[1]**; the asymmetric stretch and bending of CO_2 do so, but the only vibration of N_2 (a symmetric stretch of two identical atoms) does not **[1]**.
(c) More photosynthesis in the northern spring and summer removes CO_2; in autumn and winter, decay and reduced photosynthesis increase it **[1]**.
(d) Any two: rising sea levels (thermal expansion, melting ice); changes in crop yields; desertification; more extreme weather; loss of habitats or species **[2]**.
(e) Both could be caused by a third factor; a causal link needs a mechanism and controlled experiments (or models) **[1]**.

27. (a) $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$ **[1]**; chlorophyll **[1]**.
(b) $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$ **[1]**; yeast (enzymes) **[1]**; about 35–37 °C in the absence of oxygen **[1]**.
(c) 500/180.18 = 2.775 mol glucose → 5.55 mol ethanol **[1]**; × 46.08 = **256 g [1]**.
(d) Sugar cane can be regrown in a short time, fixing carbon from the air by photosynthesis **[1]**; the CO_2 released on burning was recently absorbed, so in principle there is no net increase **[1]**; but fossil fuels are used for fertilizers, machinery, transport and distillation, so it is not fully carbon neutral **[1]**.
(e) Any two: uses land needed for food; deforestation; high water and fertilizer use; soil depletion; lower specific energy; large volumes to transport **[2]**.

28. (a) 3 × 891 = **2673 kJ [1]**.
(b) About 95 % of the energy of burning glucose directly **[1]**; very little energy is lost in the decomposition, and the methane is a clean, easily transported gaseous fuel (but the process also releases CO_2, and leaked methane is a potent greenhouse gas) **[1]**.

29. (a) anode: $H_2 \rightarrow 2H^+ + 2e^-$; cathode: $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$ **[2]**.
(b) anode: $CH_3OH + H_2O \rightarrow CO_2 + 6H^+ + 6e^-$ **[1]**; cathode: $1½O_2 + 6H^+ + 6e^- \rightarrow 3H_2O$ **[1]**; overall: $CH_3OH + 1½O_2 \rightarrow CO_2 + 2H_2O$ **[1]**.
(c) Advantages, any two: more efficient; only water produced (no CO_2, CO, particulates or NO_x); quiet **[2]**. Disadvantage, any one: hydrogen hard to store and transport; mostly made from fossil fuels; expensive platinum catalyst **[1]**.
(d) Similarity: both use a spontaneous redox reaction to produce electricity (oxidation at anode, reduction at cathode) **[1]**. Difference: fuel cell reactants are supplied continuously; a primary cell contains a fixed amount and runs out **[1]**.

30. (a) $C_2H_5OH \rightarrow 2CO_2$: add $3H_2O$ for O **[1]**; $\mathbf{C_2H_5OH + 3H_2O \rightarrow 2CO_2 + 12H^+ + 12e^-}$ **[1]**.
(b) $\mathbf{3O_2 + 12H^+ + 12e^- \rightarrow 6H_2O}$ **[1]**; overall: $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$ **[1]**.
(c) average oxidation state of C changes from −2 to +4: 6 per C atom × 2 atoms = 12 electrons **[1]**.

31. (a) methane **−4**; methanol **−2**; glucose **0** **[1]**.
(b) methane 891/16.05 = **55.5**; methanol 726/32.05 = **22.7** **[1]**; glucose 2803/180.18 = **15.6 kJ g^{-1}** **[1]**.
(c) The general trend supports it: the more negative the oxidation state of carbon (methane −4 → octane −2.25 → glucose 0), the higher the specific energy, because a more reduced carbon has further to be oxidized to +4 **[1]**. But it is only a rough guide: methanol and ethanol both have −2 yet differ greatly (22.7 and 29.7), because specific energy is per gram and also depends on the fraction of the mass that is oxygen, which adds mass but releases no energy **[1]**.

Where this fits in Reactivity 1

- **Booklet 1** — Reactivity 1.1: measuring enthalpy changes.
- **Booklets 2 and 3** — Reactivity 1.2: bond enthalpies, Hess's law, and (HL) ΔH_f, ΔH_c and Born–Haber cycles.
- **Booklet 4** (this booklet) — Reactivity 1.3: energy from fuels.
- **Next:** HL Reactivity 1.4 — entropy and spontaneity ($\Delta G = \Delta H - T\Delta S$).

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