Chemistry Units 3 & 4

Chapter 4 — Fuel cells

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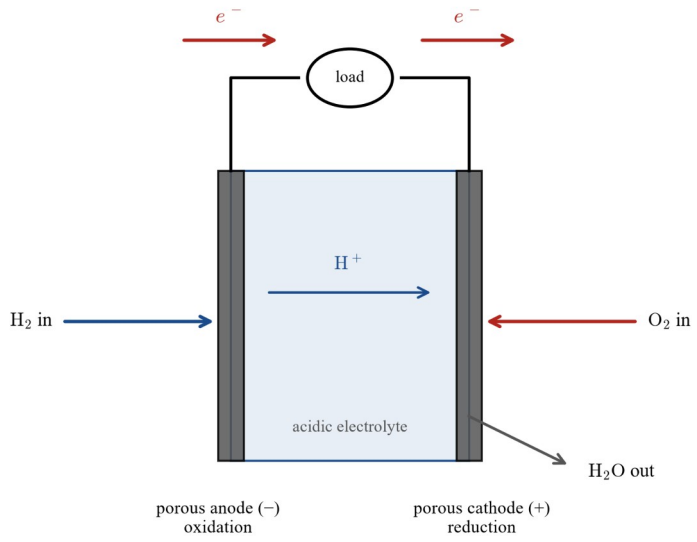
Theory

A **fuel cell** is a galvanic cell that converts the chemical energy of a fuel **directly into electrical energy** through a spontaneous redox reaction. Unlike an ordinary primary cell, its reactants are not sealed inside the cell: the fuel and an oxidant (usually oxygen from the air) are supplied continuously from outside, so the cell keeps producing electricity for as long as the supply is maintained and never "goes flat".

Direct energy conversion. In direct combustion, a fuel's chemical energy is first released as **heat**; to obtain electricity that heat must then drive a turbine and generator, so the chemical energy passes through a chain — chemical → heat → mechanical → electrical — and a large fraction is lost as waste heat at every step. A fuel cell instead makes the fuel react in a controlled redox process in which electrons are pushed through an external circuit, so chemical energy is transformed **chemical → electrical** in a single step, with no flame. Because it avoids the heat-engine stage, a fuel cell can convert a **greater proportion** of the fuel's chemical energy into useful electrical energy than combustion can.

Common design features and operating principles. All fuel cells share the same general design (the details of any particular commercial cell are not required):

- two **electrodes** separated by an **electrolyte** that conducts ions but not electrons;
- the **fuel** is fed to one electrode and the **oxidant** to the other;
- at the **anode** the fuel is **oxidised** — this is the **negative** electrode, because it releases electrons;
- at the **cathode** the oxidant is **reduced** — this is the **positive** electrode;
- **electrons** flow from the anode to the cathode through the external circuit, driving a load (motor, lamp, etc.);
- **ions** migrate through the electrolyte to complete the circuit, and the product is removed as it forms.



Porous electrodes. The electrodes of a fuel cell are made **porous** so that the gaseous reactants can pass through them and meet the electrolyte. A porous electrode has a very large internal **surface area**, and it brings the three things needed for the electrode reaction — the **gas**, the **electrolyte** and the solid **electrode** (often coated with a catalyst) — into contact over that whole area. This greatly increases the rate of the electrode reaction and the current the cell can deliver, and so increases the cell's efficiency. A flat, non-porous electrode would offer far less contact area and a much smaller current.

Writing the electrode half-equations. The reaction at each electrode is written as a balanced half-equation, including states, with the electrons shown explicitly; the two half-equations are then scaled so their electrons match and added to give the **overall** cell reaction (the electrons cancel). For the **hydrogen–oxygen** fuel cell drawn above, in an **acidic** electrolyte:

$$\text{anode (oxidation)}: \mathrm{H_2(g) \rightarrow 2H^+(aq) + 2e^-}$$

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

Multiplying the anode equation by two and adding gives the overall reaction, with the H^+ ions and electrons cancelling:

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

The **same overall reaction** occurs in a cell with an **alkaline** electrolyte, but the half-equations are written using OH^- and H_2O rather than H^+:

$$\text{anode: } H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$$

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

Whichever electrolyte is used, the only product of a hydrogen fuel cell is **water**. The same approach applies to a fuel cell supplied with a hydrocarbon or an alcohol as its fuel: in an acidic electrolyte the fuel is oxidised completely to CO_2 at the anode, releasing H^+ and electrons, while O_2 is reduced at the cathode.

Comparison with a galvanic cell and with combustion. A fuel cell is a kind of galvanic cell, so it shares the same core features — a spontaneous redox reaction, a negative anode where oxidation occurs, a positive cathode where reduction occurs, electron flow through the external circuit and ion flow through the electrolyte. The important differences are:

- **Fuel cell vs a primary (galvanic) cell:** in a primary cell the reactants are stored inside and are used up, so the cell eventually goes flat and is discarded. In a fuel cell the reactants are fed in continuously from external tanks, so it operates without interruption while fuel is supplied and does not run down. A fuel cell is **not** a rechargeable (secondary) cell either — it is **refuelled**, not recharged, so its electrode processes are not reversed by an external power supply.
- **Fuel cell vs direct combustion:** burning the fuel converts its chemical energy first to heat, which must then be turned into electricity through a heat engine, wasting much of the energy; the fuel cell converts chemical energy to electricity **directly**, so it reaches a **higher efficiency** and, with no flame, tends to produce fewer pollutants such as oxides of nitrogen and soot.

Innovation and green chemistry. Fuel cells are a focus of contemporary innovation aimed at meeting society's energy needs sustainably, guided by the green chemistry principles of **designing for energy efficiency** and **using renewable feedstocks**. Because a fuel cell converts chemical energy to electricity directly, it wastes less energy than combustion, and its low-grade waste heat can be captured and used (cogeneration) — an application of *design for energy efficiency*. The choice of fuel matters too: the hydrogen for a hydrogen fuel cell can be made either as **'grey' hydrogen**, produced from fossil natural gas (which releases CO_2), or as **'green' hydrogen**, produced by the electrolysis of water using electricity from renewable sources such as solar or wind. A fuel cell running on green hydrogen draws on a **renewable feedstock** and emits only water, so it can supply energy with very low net carbon emissions.

Worked examples

Worked Example 1 — The hydrogen–oxygen fuel cell illustrated above uses an acidic electrolyte. Write the half-equation for the process at each electrode and the overall cell reaction, and state the polarity of each electrode and the direction of electron and ion movement.

Working	Comment
$H_2(g) \rightarrow 2H^+(aq) + 2e^-$	Hydrogen is fed to the porous **anode**, where it is **oxidised** — the anode is the **negative** electrode.
$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$	Oxygen is fed to the porous **cathode**, where it is **reduced** — the cathode is the **positive** electrode.
$2\times$ anode + cathode, then cancel $4H^+$ and $4e^-$	Scale the anode to 4 electrons so the electrons and H^+ cancel on adding.

Working	Comment
$2\,H_2(g)+O_2(g)\rightarrow 2\,H_2O(l)$	Overall reaction; the only product is water.
e^-: anode → cathode (external wire); H^+: anode → cathode (electrolyte)	Matches the arrows in the diagram; water leaves at the cathode.

Worked Example 2 — The same cell can be operated with a hot **alkaline** (potassium hydroxide) electrolyte instead of an acidic one. Write the two electrode half-equations and the overall reaction for the alkaline cell, and state which way the hydroxide ions migrate.

Working	Comment
$H_2(g)+2\,OH^-(aq)\rightarrow 2\,H_2O(l)+2\,e^-$	Anode oxidation: in alkaline solution H_2 reacts with OH^- to form water.
$O_2(g)+2\,H_2O(l)+4\,e^-\rightarrow 4\,OH^-(aq)$	Cathode reduction: O_2 is reduced using water, regenerating OH^-.
$2\times$ anode + cathode; cancel $4\,OH^-$ and $2\,H_2O$	The hydroxide consumed at the anode is remade at the cathode.
$2\,H_2(g)+O_2(g)\rightarrow 2\,H_2O(l)$	Overall reaction — identical to the acidic cell.
OH^- migrates cathode → anode	Anions move through the electrolyte toward the negative anode.

Worked Example 3 — A direct methanol fuel cell oxidises liquid methanol, CH_3OH, at the anode in an acidic electrolyte, while oxygen from the air is reduced at the cathode. Write the two electrode half-equations and the overall equation, and state the energy conversion the cell carries out.

Anode — methanol is oxidised completely to carbon dioxide; balance O with water, H with H^+, and charge with electrons:

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

Cathode — oxygen is reduced:

$$O_2(g)+4\,H^+(aq)+4\,e^-\rightarrow 2\,H_2O(l)$$

The lowest common multiple of the electrons is 12, so take $2\times$ anode and $3\times$ cathode and add; the electrons and the surplus H^+ and H_2O cancel:

$$2\,CH_3OH(l)+3\,O_2(g)\rightarrow 2\,CO_2(g)+4\,H_2O(l)$$

∴ the cell converts the chemical energy of methanol **directly into electrical energy** (with some energy released as heat), oxidising the fuel to carbon dioxide and water without a flame.

Worked Example 4 — Explain why a hydrogen fuel cell can convert a larger fraction of hydrogen's chemical energy into useful electrical energy than a power station that burns the same hydrogen, and give one further way in which a fuel cell differs from an ordinary primary galvanic cell.

When hydrogen is **burned**, its chemical energy is first released as **heat**; to generate electricity that heat must raise steam to spin a turbine and generator, so the energy passes through a chemical → heat → mechanical → electrical chain, and a large fraction is lost as waste heat at each stage, limited by the efficiency of the heat engine. A **fuel cell** instead makes the hydrogen react in a controlled redox process in which the electrons are forced through the external circuit, converting chemical energy **directly into electrical energy** in a single step with no combustion. Because it bypasses the wasteful heat-engine stage, a greater proportion of the chemical energy becomes useful electricity.

A further difference from a primary galvanic cell is the **supply of reactants**: in a primary cell the reactants are stored inside the cell and are gradually used up, so the cell goes flat and must be discarded, whereas a fuel cell is fed with fuel and oxidant continuously from external tanks, so it keeps operating for as long as it is supplied.

∴ the fuel cell is more efficient because it converts chemical energy directly to electricity, and it runs continuously because its reactants are supplied from outside rather than stored inside.

Practice questions

Question 1 — A hydrogen–oxygen fuel cell operates with an acidic electrolyte.

a. Write the balanced half-equation, including states, for the reaction at the anode.
b. Write the balanced half-equation, including states, for the reaction at the cathode.
c. Write the balanced overall equation for the cell reaction.

Question 2 — A hydrogen–oxygen fuel cell operates with an alkaline (KOH) electrolyte.

a. Write the balanced half-equation, including states, for the reaction at the anode.
b. Write the balanced half-equation, including states, for the reaction at the cathode.
c. Write the overall equation and state which electrode is the negative terminal of the cell.

Question 3 — Refer to the diagram of the hydrogen–oxygen fuel cell in the Theory section.

a. Name the electrode to which the hydrogen is supplied, and state its polarity (positive or negative).
b. State the direction in which electrons flow through the external circuit, and the direction in which $\mathrm{H^+}$ ions migrate through the electrolyte.
c. Identify the product of the cell and the electrode at which it forms.

Question 4 — A fuel cell is operated using methane, $\mathrm{CH_4}$, as the fuel and oxygen as the oxidant, with an acidic electrolyte.

a. Write the balanced half-equation, including states, for the oxidation of methane at the anode.
b. Write the balanced half-equation, including states, for the reaction at the cathode.
c. Write the balanced overall equation for the cell reaction.

Question 5 — Consider the general operation of a fuel cell.

a. State the energy conversion that a fuel cell carries out.
b. Explain the role of the porous electrodes in a fuel cell.
c. Explain why a fuel cell converts a greater fraction of a fuel's chemical energy into useful energy than direct combustion of the fuel does.

Question 6 — Compare a fuel cell with other ways of obtaining energy from a fuel.

a. State one feature that a fuel cell shares with a primary galvanic cell.
b. State two ways in which a fuel cell differs from a primary galvanic cell.
c. State one advantage of a fuel cell over the direct combustion of the fuel.

Question 7 — A fuel cell uses ethanol, $\mathrm{C_2H_5OH}$, as the fuel with an acidic electrolyte and oxygen as the oxidant. Write the balanced half-equation for the anode, the balanced half-equation for the cathode, and the balanced overall equation, all including states.

Question 8 — A fuel cell is run on propane, $\mathrm{C_3H_8}$, with an acidic electrolyte. Write the anode half-equation, the cathode half-equation and the overall equation, including states.

Question 9 — In an ethane fuel cell with an acidic electrolyte, the two electrode processes are

$$\mathrm{C_2H_6(g) + 4\,H_2O(l) \rightarrow 2\,CO_2(g) + 14\,H^+(aq) + 14\,e^-}$$

$$\mathrm{O_2(g) + 4\,H^+(aq) + 4\,e^- \rightarrow 2\,H_2O(l)}$$

Deduce the balanced overall equation for the cell, and state which of the two half-equations occurs at the negative electrode.

Question 10 — Explain the energy transformations that occur in a hydrogen fuel cell, and explain why the cell achieves a higher efficiency than a petrol engine used to do the same work. Set out your answer as a clear reasoning chain.

Question 11 — Explain how the use of **porous** electrodes increases the efficiency of a fuel cell. In your answer refer to surface area and to the contact between the gas, the electrolyte and the electrode.

Question 12 — Compare a hydrogen fuel cell with the direct combustion of hydrogen gas. Give one similarity and two differences, and state which process converts a greater fraction of the chemical energy into useful work.

Question 13 — A student building a model hydrogen–oxygen fuel cell with an acidic electrolyte suggests replacing the electrolyte with a strip of copper metal, arguing that "copper conducts electricity far better than any solution, so the cell will deliver a larger current". Evaluate the student's suggestion. In your answer, state the function of the electrolyte in a fuel cell and explain what would happen to the electrons and to the ions if the change were made.

Question 14 — In a hydrogen–oxygen fuel cell:

a. state which electrode is the positive electrode;
b. state the direction in which electrons flow through the external load, and explain why they flow in that direction;
c. state at which electrode the oxygen reacts, and whether the oxygen is oxidised or reduced.

Question 15 — A hydrogen–oxygen fuel cell with an acidic electrolyte is powered by 'green' hydrogen produced by the electrolysis of water using solar electricity, in place of 'grey' hydrogen produced from natural gas. Write the anode half-equation, the cathode half-equation and the overall equation (including states), and explain, with reference to the green chemistry principle of using renewable feedstocks, why the switch to green hydrogen makes the cell a more sustainable energy source.

Question 16 — Explain why a hydrogen–oxygen fuel cell never needs to be recharged and can deliver current for as long as it is supplied with reactants. Contrast this with the operation of a rechargeable (secondary) cell, using correct terminology for what happens to the electrode reactions.

Question 17 — A fuel cell is run on propan-1-ol, C_3H_7OH, with an acidic electrolyte and oxygen as the oxidant. Write the anode half-equation, the cathode half-equation and the balanced overall equation, all including states.

Question 18 — A city council replaces its fleet of diesel buses with buses powered by hydrogen fuel cells. The hydrogen is 'green' hydrogen, produced by the electrolysis of water using electricity from a solar farm. Evaluate the fuel-cell buses against the diesel buses with reference to **two** green chemistry principles. In your response:

- name and apply the green chemistry principle of **design for energy efficiency**, supported by evidence;
- name and apply the green chemistry principle of **use of renewable feedstocks**, supported by evidence;
- give a concluding statement about which option is the more sustainable.

Question 19 — A car manufacturer offers the same model with either a hydrogen fuel-cell drivetrain or a petrol internal-combustion engine. Discuss, with reference to the energy conversions involved, the efficiency of each, and the products released, why the fuel-cell version can be described as the cleaner and more efficient option. Identify one limitation of the fuel-cell option, and give a concluding statement.

Answers

1 a $H_2(g) \rightarrow 2H^+(aq) + 2e^-$; b $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$; c $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$. **2** a $H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$; b $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$; c $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$; the **anode** (where H_2 is oxidised) is the negative terminal. **3** a the **anode**, which is the **negative** electrode; b electrons flow anode → cathode through the external circuit; H^+ ions migrate anode → cathode through the electrolyte; c the product is **water**, formed at the **cathode**. **4** a $CH_4(g) + 2H_2O(l) \rightarrow CO_2(g) + 8H^+(aq) + 8e^-$; b $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$; c $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$. **5** a chemical energy → electrical energy (directly); b porous electrodes give a large surface area and bring the gas, electrolyte and electrode into contact, speeding the electrode reactions and increasing the current/efficiency; c the fuel cell converts chemical energy directly to electricity, avoiding the lossy heat → mechanical → electrical steps of combustion. **6** a e.g. both are galvanic cells using a spontaneous redox reaction, with a negative anode (oxidation) and a positive cathode (reduction); b e.g. the fuel cell is supplied with reactants continuously from outside (a primary cell stores them inside), and it does not go flat / is refuelled rather than discarded; c e.g. higher efficiency (direct conversion) and/or fewer pollutants than combustion. **7** anode $C_2H_5OH(l) + 3H_2O(l) \rightarrow 2CO_2(g) + 12H^+(aq) + 12e^-$; cathode $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$; overall $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$. **8** anode $C_3H_8(g) + 6H_2O(l) \rightarrow 3CO_2(g) + 20H^+(aq) + 20e^-$; cathode $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$; overall $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$. **9** $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$; the ethane oxidation half-equation occurs at the negative electrode (the anode). **10** chemical → electrical directly (single redox step, no combustion), so less energy is lost as waste heat than in a petrol engine, which goes chemical → heat → mechanical and is limited by heat-engine efficiency; hence higher efficiency. **11** porous electrodes have a large surface area and bring gas, electrolyte and (catalytic) electrode into three-way contact, so the electrode reactions are faster and the cell delivers more current — greater efficiency. **12** similarity: both are exothermic redox reactions with the same overall equation $2H_2 + O_2 \rightarrow 2H_2O$; differences: the cell produces electricity directly / via a controlled redox at electrodes with no flame, whereas combustion releases the energy as heat and light; the cell converts a greater fraction of the chemical energy into useful work. **13** the suggestion fails: the electrolyte must conduct **ions but not electrons** — it carries H^+ from anode to cathode to complete the circuit while forcing the electrons through the external circuit; copper conducts electrons, so it would short-circuit the cell internally (electrons would pass straight from anode to cathode inside the cell, no current would reach the load and the energy would be released as heat), and a solid metal cannot supply or transport the $H^+(aq)$ the half-equations require, so the electrode reactions would stop. **14** a the cathode; b electrons flow from the anode (negative) to the cathode (positive) through the load, because oxidation of the hydrogen at the anode releases electrons while reduction of the oxygen at the cathode consumes them; c oxygen reacts at the cathode and is **reduced**. **15** equations as in Q1; the cell is more sustainable because green hydrogen is made from water using renewable (solar) electricity — a renewable feedstock that is replenished naturally — rather than from finite natural gas, and its production and use add very little net CO_2. **16** the reactants are supplied continuously from outside, so the cell does not use up a stored charge and never needs recharging; a secondary cell stores its reactants internally and must have its electrode reactions **reversed** by an external power supply to be recharged. **17** anode $C_3H_7OH(l) + 5H_2O(l) \rightarrow 3CO_2(g) + 18H^+(aq) + 18e^-$; cathode $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$; overall $2C_3H_7OH(l) + 9O_2(g) \rightarrow 6CO_2(g) + 8H_2O(l)$. **18** evaluation (see solutions): design for energy efficiency — the fuel cell converts chemical energy directly to electricity, wasting less than a diesel engine; use of renewable feedstocks — green hydrogen from solar-powered electrolysis of water is renewable, unlike diesel from crude oil; conclusion — the fuel-cell buses are the more sustainable option. **19** the fuel cell converts chemical → electrical directly (higher efficiency) and emits only water, whereas the petrol engine goes chemical → heat → mechanical (lower efficiency) and emits CO_2 and pollutants; limitation e.g. hydrogen storage/refuelling infrastructure or the source of the hydrogen; conclusion — the fuel-cell version is cleaner and more efficient.

Fully-worked solutions

Question 1 a. Hydrogen is oxidised at the anode: $H_2(g) \rightarrow 2H^+(aq) + 2e^-$ (2 H and zero net charge on each side). b. Oxygen is reduced at the cathode: $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$ (O and H balance; charge $= 0$ on each side).

c. Doubling the anode equation to match 4 electrons and adding cancels the H^+ and electrons: $2H_2(g)+O_2(g) \rightarrow 2H_2O(l)$.

Question 2 a. In alkaline solution the hydrogen is oxidised using hydroxide ions: $H_2(g)+2OH^-(aq) \rightarrow 2H_2O(l)+2e^-$ (2 O, 4 H and charge -2 on each side). b. Oxygen is reduced to hydroxide: $O_2(g)+2H_2O(l)+4e^- \rightarrow 4OH^-(aq)$ (4 O, 4 H and charge -4 on each side). c. Doubling the anode equation and adding cancels $4OH^-$ and $2H_2O$, giving $2H_2(g)+O_2(g) \rightarrow 2H_2O(l)$. The anode, where oxidation occurs and electrons leave the cell, is the **negative** terminal.

Question 3 a. Hydrogen (the fuel) is supplied to the **anode**, which in a galvanic-type cell is the **negative** electrode because oxidation there releases electrons. b. Electrons flow from the anode to the cathode through the external circuit; H^+ ions migrate in the same overall direction — from the anode to the cathode — through the electrolyte, completing the circuit. c. The product is **water**; it forms at the **cathode**, where oxygen combines with H^+ and electrons.

Question 4 a. Balance C with CO_2, O with water, H with H^+, then charge with electrons: $CH_4(g)+2H_2O(l) \rightarrow CO_2(g)+8H^+(aq)+8e^-$ (C 1, H 8, O 2 and charge 0 on each side). b. $O_2(g)+4H^+(aq)+4e^- \rightarrow 2H_2O(l)$. c. Take $2\times$ cathode to match 8 electrons and add; the $8H^+$, the electrons and 4 of the 6 water molecules cancel: $CH_4(g)+2O_2(g) \rightarrow CO_2(g)+2H_2O(l)$.

Question 5 a. A fuel cell converts **chemical energy directly into electrical energy**. b. The porous electrodes provide a very large surface area and allow the gaseous reactants to pass through and meet the electrolyte, bringing the gas, the electrolyte and the solid (often catalytic) electrode into contact over that whole area. This increases the rate of the electrode reactions and the current the cell can supply, raising its efficiency. c. In direct combustion the chemical energy is released first as heat, which must then be converted to electricity through a heat engine (chemical $\rightarrow$ heat $\rightarrow$ mechanical $\rightarrow$ electrical), losing a large fraction as waste heat at each stage and being limited by the heat-engine efficiency. A fuel cell converts chemical energy straight to electricity in one redox step, so a greater fraction of the energy is captured as useful work.

Question 6 a. Both are galvanic cells that generate electricity from a **spontaneous** redox reaction, with a negative anode where oxidation occurs and a positive cathode where reduction occurs. b. (1) A fuel cell is supplied with its reactants continuously from external tanks, whereas a primary cell has its reactants stored inside it. (2) A fuel cell keeps operating for as long as fuel is supplied and does not go flat, whereas a primary cell runs down as its stored reactants are consumed and is then discarded. c. A fuel cell converts chemical energy to electricity directly, so it is more efficient than combustion (and produces fewer pollutants because there is no flame).

Question 7 Anode — oxidise ethanol fully to CO_2, balancing O with water, H with H^+ and charge with electrons:

$$C_2H_5OH(l)+3H_2O(l) \rightarrow 2CO_2(g)+12H^+(aq)+12e^-$$

(C 2, H 12, O 4 and charge 0 on each side.) Cathode: $O_2(g)+4H^+(aq)+4e^- \rightarrow 2H_2O(l)$. Overall — take $3\times$ cathode to match 12 electrons and add; the $12H^+$, the electrons and 3 water molecules cancel:

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

Question 8 Anode: $C_3H_8(g)+6H_2O(l) \rightarrow 3CO_2(g)+20H^+(aq)+20e^-$ (C 3, H 20, O 6, charge 0 each side). Cathode: $O_2(g)+4H^+(aq)+4e^- \rightarrow 2H_2O(l)$. Overall — take $5\times$ cathode to match 20 electrons and add; $20H^+$, the electrons and 6 water molecules cancel:

$$C_3H_8(g)+5O_2(g) \rightarrow 3CO_2(g)+4H_2O(l)$$

Question 9 The lowest common multiple of the electrons (14 and 4) is 28: take $2\times$ the ethane half-equation (28 electrons) and $7\times$ the oxygen half-equation (28 electrons) and add. The 28 electrons and $28H^+$ cancel, and 8 of the 14 water molecules cancel:

$$2\,C_2H_6(g)+7\,O_2(g)\rightarrow 4\,CO_2(g)+6\,H_2O(l)$$

(C 4, H 12, O 14 on each side.) The ethane oxidation half-equation, in which electrons are released, occurs at the **negative electrode** (the anode).

Question 10 In a hydrogen fuel cell the fuel undergoes a controlled redox reaction at the electrodes: hydrogen is oxidised at the anode and oxygen is reduced at the cathode, and the electrons released are forced through the external circuit. The energy transformation is therefore **chemical → electrical**, achieved directly in a single step with no combustion. A petrol engine instead **burns** its fuel, releasing the chemical energy first as **heat**, which is then turned into the mechanical motion of the pistons (chemical → heat → mechanical). Each of these conversions wastes a large fraction of the energy as heat to the exhaust and cooling system, and the maximum possible efficiency is limited by the operating temperatures of the heat engine. Because the fuel cell skips the heat and mechanical stages and converts chemical energy straight into electricity, less energy is lost and a higher fraction becomes useful work — so the fuel cell is more efficient.

Question 11 A porous electrode is riddled with tiny channels, giving it an internal surface area far larger than its outer dimensions suggest. The gaseous reactant can diffuse into these channels so that the **gas**, the liquid **electrolyte** and the solid **electrode** (usually coated with a catalyst) all meet along an enormous three-way contact area. The electrode reaction can only take place where all three are present, so this large contact area allows many more reactant particles to react per second. The rate of the electrode reactions and the current the cell can deliver therefore increase, which raises the cell's efficiency compared with a flat, non-porous electrode.

Question 12 Similarity: both processes are exothermic redox reactions between hydrogen and oxygen with the same overall equation, $2\,H_2(g)+O_2(g)\rightarrow 2\,H_2O(l)$, and both are spontaneous. Differences: (1) in the fuel cell the oxidation and reduction happen at **separate electrodes** so that electrons travel through an external circuit as an electric current, whereas in combustion the electrons transfer directly between the reacting molecules and the energy is released as **heat and light** (a flame); (2) the fuel cell delivers the energy as **electricity** at a comparatively low temperature, whereas combustion delivers it as high-temperature heat. The **fuel cell** converts a greater fraction of the chemical energy into useful work, because it produces electricity directly instead of via the lossy heat-to-work conversion of a heat engine.

Question 13 The electrolyte does two jobs in a fuel cell. First, it supplies and transports the ions that appear in the half-equations: the $H^+(aq)$ released at the anode, $H_2(g)\rightarrow 2\,H^+(aq)+2\,e^-$, migrates through the electrolyte to the cathode, where it is consumed, $O_2(g)+4\,H^+(aq)+4\,e^-\rightarrow 2\,H_2O(l)$. This internal ion flow completes the circuit and stops charge building up at the electrodes. Second — and just as important — the electrolyte must **not** conduct electrons: it is only because electrons cannot cross it that they are forced to travel from the anode to the cathode through the external circuit, where they do useful work.

Copper is a metallic conductor, so it carries electrons. Replacing the electrolyte with a copper strip would join the two electrodes internally and **short-circuit** the cell: the electrons released at the anode would take the direct path through the copper to the cathode instead of through the load, so no current would flow in the external circuit and the cell would deliver no electrical energy — the energy of the reaction would simply be released as heat. A solid metal also cannot supply or transport the $H^+(aq)$ ions that both half-equations require, so the electrode reactions could not be sustained.

The student's suggestion is therefore incorrect. A fuel cell needs an electrolyte that conducts ions but not electrons; copper, being a good conductor of electrons, would short-circuit the cell rather than increase the current it delivers.

Question 14 a. The **cathode** is the positive electrode (reduction of oxygen occurs there). b. Electrons flow from the **anode** (negative) to the **cathode** (positive) through the external load. Oxidation of the hydrogen at the anode releases electrons into that electrode, while reduction of the oxygen at the cathode removes them, so electrons are driven along the external circuit from anode to cathode. c. Oxygen reacts at the **cathode** and is **reduced** (it gains electrons: $O_2+4\,H^++4\,e^-\rightarrow 2\,H_2O$).

Question 15 The electrode reactions are those of the acidic hydrogen–oxygen cell:

$$\text{anode: } H_2(g)\rightarrow 2\,H^+(aq)+2\,e^-$$

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

$$\text{overall: } 2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

The fuel cell itself works the same way whichever hydrogen it is supplied with, but the **feedstock** differs. 'Grey' hydrogen is made from **natural gas**, a finite fossil resource, and its manufacture releases carbon dioxide. 'Green' hydrogen is made by the **electrolysis of water using renewable (solar) electricity**: the feedstock is water, which is renewable, and the energy comes from the Sun, which is continually replenished. Using a renewable feedstock in place of a fossil one means the overall process adds very little net CO_2 to the atmosphere — the cell still emits only water — so switching to green hydrogen makes the energy supply far more sustainable, in line with the green chemistry principle of using renewable feedstocks.

Question 16 A fuel cell does not rely on a fixed amount of stored reactant: its fuel and oxidant are fed in continuously from outside, so it can keep delivering current for as long as the supply is maintained and it never accumulates a "used-up" state that would need to be reversed. When its fuel tank is empty it is simply **refuelled**. A rechargeable (secondary) cell is different: it stores its reactants internally, and as it discharges those reactants are converted to products. To recharge it, an external power supply drives the cell in reverse so that the **electrode reactions are reversed**, regenerating the original reactants (the cell operates as an electrolytic cell during charging). A fuel cell is refuelled rather than recharged, so its electrode reactions are never reversed.

Question 17 Anode — oxidise propan-1-ol fully to CO_2: $C_3H_7OH(l) + 5H_2O(l) \rightarrow 3CO_2(g) + 18H^+(aq) + 18e^-$ (C 3, H 18, O 6, charge 0 each side). Cathode: $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$. Overall — the lowest common multiple of the electrons (18 and 4) is 36: take $2\times$ anode and $9\times$ cathode and add; the $36H^+$, electrons and 10 of the 18 water molecules cancel:

$$2C_3H_7OH(l) + 9O_2(g) \rightarrow 6CO_2(g) + 8H_2O(l)$$

Question 18 Design for energy efficiency. A hydrogen fuel cell converts the chemical energy of its fuel **directly into electrical energy** to drive the bus's motor, whereas a diesel engine burns its fuel and converts chemical energy to heat and then to mechanical work, wasting much of it as heat in the exhaust and cooling system. Because the fuel cell avoids the heat-engine stage it captures a greater fraction of the fuel's energy as useful work, and the low-grade waste heat it does produce can be recovered — evidence that the fuel-cell design meets the principle of designing for energy efficiency far better than the diesel engine.

Use of renewable feedstocks. The fuel-cell buses run on **green hydrogen** made by electrolysing water with solar electricity — the feedstock (water) is renewable and the energy source (sunlight) is continually replenished, and the only product when the hydrogen is used is water. The diesel buses run on **diesel refined from crude oil**, a finite fossil resource that cannot be replaced within a human timescale, and their combustion releases CO_2 and pollutants. The fuel-cell buses therefore satisfy the principle of using renewable feedstocks, whereas the diesel buses do not.

Conclusion. Judged against both design for energy efficiency and the use of renewable feedstocks, the hydrogen fuel-cell buses running on green hydrogen are clearly the **more sustainable** option; the main practical trade-offs are the need to build hydrogen production, storage and refuelling infrastructure and to ensure the electricity used really is renewable.

Question 19 In the fuel-cell car, hydrogen and oxygen react in a controlled redox reaction so that chemical energy is converted **directly into electrical energy**, which then drives an electric motor; the only product is **water**. In the petrol car the fuel is **burned**, converting chemical energy first into heat and then into the mechanical motion of the engine (chemical → heat → mechanical), and it emits **carbon dioxide** along with pollutants such as oxides of nitrogen and unburnt hydrocarbons. Because the fuel cell converts energy directly and skips the wasteful heat-engine stage, it is **more efficient** — a greater fraction of the fuel's energy reaches the wheels — and because it emits only water at the point of use it is **cleaner**. One limitation of the fuel-cell option is the difficulty and cost of producing, storing and distributing hydrogen (and the fact that it is only truly clean if the hydrogen is 'green'). Concluding statement: provided the hydrogen is produced renewably, the fuel-cell version is both cleaner and more efficient than the petrol version, and is the better choice on environmental grounds.

Chapter 4 Fuel cells — 4B Cell stoichiometry and Faraday's laws

Theory

A galvanic cell and a fuel cell both release electrical energy from a spontaneous redox reaction: electrons produced by oxidation at the anode travel through the external circuit to the cathode, where reduction occurs. **Faraday's laws** connect the electric charge carried by those electrons to the amounts of reactant consumed and product formed at the electrodes. This lets us calculate the mass or volume of a cell's reactants and products, and the current or time involved.

Charge, current and time. The quantity of electric charge Q that flows is the product of the current and the time for which it flows:

$$Q = I\,t$$

where Q is in **coulombs** (C) when the current I is in **amperes** (A) and the time t is in **seconds** (s). Because this relationship demands SI units, a current given in milliamperes must be converted to amperes ($1\,\text{mA} = 10^{-3}\,\text{A}$) and a time given in minutes or hours must be converted to seconds ($1\,\text{min} = 60\,\text{s}$, $1\,\text{h} = 3600\,\text{s}$) *before* substituting.

The Faraday constant and the amount of electrons. One mole of electrons carries a charge of $F = 96\,500\,\text{C mol}^{-1}$, the **Faraday constant**. The amount, in mole, of electrons that has passed is therefore

$$n(\text{e}^-) = \frac{Q}{F}$$

Linking electrons to reactants and products. The **electrode half-equation** fixes the ratio between electrons and the species reacting or forming at that electrode. For a hydrogen–oxygen fuel cell operating in acidic solution:

$$\text{anode: } \text{H}_2(\text{g}) \rightarrow 2\,\text{H}^+(\text{aq}) + 2\,\text{e}^-$$

$$\text{cathode: } \text{O}_2(\text{g}) + 4\,\text{H}^+(\text{aq}) + 4\,\text{e}^- \rightarrow 2\,\text{H}_2\text{O}(\text{l})$$

so 2 mol of electrons pass for every 1 mol of H_2 oxidised, and 4 mol of electrons for every 1 mol of O_2 reduced. The same idea governs a discharging galvanic cell: a reactive metal anode is oxidised, for example $\text{Zn}(\text{s}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\,\text{e}^-$ (2 mol e⁻ per mol Zn) or $\text{Al}(\text{s}) \rightarrow \text{Al}^{3+}(\text{aq}) + 3\,\text{e}^-$ (3 mol e⁻ per mol Al), while a metal ion may be reduced and deposited at the cathode, for example $\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$ (1 mol e⁻ per mol Ag). The amount of any electrode species follows from the half-equation ratio:

$$n(\text{species}) = n(\text{e}^-) \times \frac{\text{coefficient of the species}}{\text{coefficient of e}^-}$$

From amount to mass or gas volume. Once the amount of the electrode species is known, it is converted to a **mass** with $n = \frac{m}{M}$, or, for a gas measured at **SLC** (25 °C, 100 kPa), to a **volume** with

$$V = n \times V_m,\; V_m = 24.8\,\text{L mol}^{-1}$$

The standard chain — and working backwards. Most 4B problems follow one four-step chain: (1) charge, $Q = I\,t$; (2) electrons, $n(\text{e}^-) = Q / F$; (3) electrode species, $n(\text{species})$ from the half-equation ratio; (4) mass or volume, $m = n\,M$ or $V = n\,V_m$. When the quantity of reactant or product is given instead and the **current or time** is required, the same chain is run in reverse: find $n(\text{species})$, then $n(\text{e}^-)$, then $Q = n(\text{e}^-) \times F$, and finally $I = \frac{Q}{t}$ or $t = \frac{Q}{I}$. A "how long can the cell run" problem is exactly this — the total charge available from a fixed mass of fuel sets the maximum time the cell can deliver a given current.

Discharge only. Throughout this subtopic the cell is *delivering* current — a galvanic cell or fuel cell operating spontaneously — so the anode reactant is consumed and the cathode products form as written above. Applying the

same laws to the products formed when an external supply *drives* a non-spontaneous reaction (electrolysis) is treated separately.

Units and significant figures. Every answer must carry its unit, and the number of significant figures should match the data (usually 3). Keeping the current in amperes and the time in seconds throughout — and converting milliamperes and hours first — avoids the most common Faraday's-law error.

Worked examples

Worked Example 1 — A hydrogen–oxygen fuel cell delivers a steady current of 5.00 A for 30.0 min. Calculate the mass of hydrogen gas consumed at the anode. (In acidic solution the anode reaction is $\mathrm{H_2(g) \rightarrow 2\,H^+(aq) + 2\,e^-}$.)

Working	Comment
$t = 30.0\,\text{min} = 30.0 \times 60 = 1800\,\text{s}$	Convert the time to seconds before using $Q = I\,t$.
$Q = I\,t = 5.00 \times 1800 = 9000\,\text{C}$	Charge delivered by the cell.
$n(\mathrm{e^-}) = \dfrac{Q}{F} = \dfrac{9000}{96\,500} = 0.09326\,\text{mol}$	Amount of electrons; $F = 96\,500\,\text{C mol}^{-1}$.
$n(\mathrm{H_2}) = \dfrac{n(\mathrm{e^-})}{2} = \dfrac{0.09326}{2} = 0.04663\,\text{mol}$	The anode half-equation gives 2 mol $\mathrm{e^-}$ per mol $\mathrm{H_2}$.
$m(\mathrm{H_2}) = n \times M = 0.04663 \times 2.0 = 0.0933\,\text{g}$	Mass of hydrogen consumed (3 sig figs).

Worked Example 2 — A galvanic cell has a magnesium anode, $\mathrm{Mg(s) \rightarrow Mg^{2+}(aq) + 2\,e^-}$. Over 25.0 min the magnesium electrode loses 0.300 g. Calculate the average current delivered by the cell.

Working	Comment
$n(\mathrm{Mg}) = \dfrac{m}{M} = \dfrac{0.300}{24.3} = 0.012346\,\text{mol}$	Amount of magnesium oxidised at the anode.
$n(\mathrm{e^-}) = 2 \times 0.012346 = 0.024691\,\text{mol}$	2 mol $\mathrm{e^-}$ per mol Mg, from the half-equation.
$Q = n(\mathrm{e^-}) \times F = 0.024691 \times 96\,500 = 2382.7\,\text{C}$	Charge that flowed (rearranging $n(\mathrm{e^-}) = Q / F$).
$t = 25.0\,\text{min} = 1500\,\text{s}$	Convert the time to seconds.
$I = \dfrac{Q}{t} = \dfrac{2382.7}{1500} = 1.59\,\text{A}$	Average current (3 sig figs).

Worked Example 3 — A hydrogen–oxygen fuel cell operates at a steady current of 2.00 A for 45.0 min. Calculate the volume of oxygen gas, measured at SLC, consumed at the cathode.

The cathode reaction is $\mathrm{O_2(g) + 4\,H^+(aq) + 4\,e^- \rightarrow 2\,H_2O(l)}$, giving 4 mol $\mathrm{e^-}$ per mol $\mathrm{O_2}$.

$$t = 45.0\,\text{min} = 2700\,\text{s}, Q = I\,t = 2.00 \times 2700 = 5400\,\text{C}$$

$$n(\mathrm{e^-}) = \frac{5400}{96\,500} = 0.055959\,\text{mol}$$

$$n(\mathrm{O_2}) = \frac{n(\mathrm{e^-})}{4} = \frac{0.055959}{4} = 0.013990\,\text{mol}$$

$$V(\mathrm{O_2}) = n \times V_m = 0.013990 \times 24.8 = 0.347\,\text{L}$$

$\therefore$ the cell consumes $0.347\,\text{L}$ of oxygen at SLC.

Worked Example 4 — A direct methanol fuel cell oxidises methanol at the anode according to $\mathrm{CH_3OH(l) + H_2O(l) \rightarrow CO_2(g) + 6\,H^+(aq) + 6\,e^-}$. The cell delivers a steady current of 1.50 A. Calculate the time, in hours, for which it can operate on 3.20 g of methanol ($M = 32.0\,\text{g mol}^{-1}$).

$$n(\mathrm{CH_3OH}) = \frac{m}{M} = \frac{3.20}{32.0} = 0.100\,\mathrm{mol}$$

The half-equation gives 6 mol e^- per mol methanol, so

$$n(\mathrm{e^-}) = 6 \times 0.100 = 0.600\,\mathrm{mol}, Q = n(\mathrm{e^-}) \times F = 0.600 \times 96\,500 = 57\,900\,\mathrm{C}$$

$$t = \frac{Q}{I} = \frac{57\,900}{1.50} = 38\,600\,\mathrm{s} = \frac{38\,600}{3600} = 10.7\,\mathrm{h}$$

$\therefore$ the cell can run for about $10.7\,\mathrm{h}$ on 3.20 g of methanol.

Practice questions

Question 1 — A galvanic cell delivers a steady current of 3.00 A for 20.0 min.

a. Calculate the charge, in coulombs, that passes through the cell.
b. Calculate the amount, in mole, of electrons transferred.
c. The anode reaction is $\mathrm{M(s) \rightarrow M^{2+}(aq) + 2e^-}$. Calculate the amount, in mole, of metal M consumed.

Question 2 — A hydrogen–oxygen fuel cell delivers a current of 4.00 A for 12.0 min.

a. Write the balanced half-equation (with states) for the oxidation of hydrogen at the anode in acidic solution.
b. Calculate the amount, in mole, of electrons transferred.
c. Calculate the mass of hydrogen gas consumed at the anode.

Question 3 — A galvanic cell deposits copper at its cathode: $\mathrm{Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)}$. The cell delivers a current of 1.20 A.

a. Calculate the amount, in mole, of copper in a 0.500 g deposit.
b. Calculate the charge required to deposit this copper.
c. Calculate the time required, in seconds and in minutes.

Question 4 — A hydrogen–oxygen fuel cell delivers a current of 6.00 A for 20.0 min.

a. Calculate the charge that passes.
b. Calculate the amount, in mole, of electrons transferred.
c. Calculate the volume of hydrogen gas, measured at SLC, consumed at the anode.

Question 5 — In a galvanic cell the nickel anode reaction is $\mathrm{Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-}$. Over 40.0 min the nickel electrode loses 0.220 g.

a. Calculate the amount, in mole, of electrons transferred.
b. Calculate the charge that passed.
c. Calculate the average current delivered by the cell.

Question 6 — A galvanic cell deposits silver at its cathode: $\mathrm{Ag^+(aq) + e^- \rightarrow Ag(s)}$. The cell delivers a current of 250 mA for 1.50 hours.

a. Convert the current to amperes and the time to seconds.
b. Calculate the charge that passes.
c. Calculate the mass of silver deposited.

Question 7 — A fuel cell delivers a steady current of 8.00 A for 45.0 min. Calculate the charge that passes and the amount, in mole, of electrons transferred.

Question 8 — A galvanic cell has an aluminium anode, $\mathrm{Al(s) \rightarrow Al^{3+}(aq) + 3e^-}$. The cell delivers a current of 2.00 A for 50.0 min. Calculate the mass of aluminium consumed.

Question 9 — A hydrogen–oxygen fuel cell operates at 3.20 A for 25.0 min. Calculate the volume of oxygen gas, measured at SLC, consumed at the cathode.

Question 10 — A galvanic cell deposits 2.00 g of silver ($Ag^+(aq) + e^- \rightarrow Ag(s)$) at its cathode in 30.0 min. Calculate the average current delivered.

Question 11 — A hydrogen–oxygen fuel cell is required to consume 0.100 g of hydrogen gas at the anode. If it operates at a current of 1.80 A, calculate the time required, in minutes.

Question 12 — A hydrogen–oxygen fuel cell is charged with 5.00 g of hydrogen gas and powers a device that draws a constant current of 10.0 A. Calculate the length of time, in hours, for which the device can be powered before the hydrogen is exhausted.

Question 13 — A hydrogen–oxygen fuel cell consumes hydrogen at a steady rate of 0.0100 g per minute. Calculate the current, in amperes, that the cell is delivering.

Question 14 — A direct methanol fuel cell oxidises methanol at the anode: $CH_3OH(l) + H_2O(l) \rightarrow CO_2(g) + 6H^+(aq) + 6e^-$. The cell delivers 5.00 A for 1.00 hour. Calculate the volume of carbon dioxide, measured at SLC, produced at the anode.

Question 15 — A hydrogen–oxygen fuel cell operates with an alkaline electrolyte. Write balanced half-equations (with states) for the reactions at the anode and at the cathode, and the overall cell equation.

Question 16 — A hydrogen–oxygen fuel cell consumes 0.992 L of oxygen gas, measured at SLC, at its cathode over 15.0 min. Calculate the average current delivered by the cell.

Question 17 — A galvanic cell can deposit either silver ($Ag^+(aq) + e^- \rightarrow Ag(s)$) or copper ($Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$) at its cathode. Explain why depositing a given amount (in mole) of copper requires twice the charge needed to deposit the same amount of silver.

Question 18 — A galvanic cell with a zinc anode, $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$, delivers a current of 500 mA. Calculate the time, in hours, required to consume 1.50 g of zinc.

Question 19 — A hydrogen–oxygen fuel cell delivers a current of 2.40 A for 2.00 hours. Calculate the mass of water produced by the cell. The overall reaction is $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$.

Question 20 — A hydrogen–oxygen fuel cell delivers a current of 12.0 A for 90.0 min. Calculate:

a. the mass of hydrogen gas consumed at the anode;
b. the volume of oxygen gas, measured at SLC, consumed at the cathode.

Question 21 — A portable hydrogen–oxygen fuel cell powers a remote sensor that draws a constant current of 150 mA. The cell is charged with 0.400 g of hydrogen gas.

a. Calculate the maximum charge, in coulombs, that the cell can deliver from this hydrogen.
b. Calculate the maximum time, in hours, for which the sensor can be powered.
c. State one assumption made in this calculation.

Answers

1 a 3600 C; b 0.0373 mol; c 0.0187 mol. **2** a $H_2(g) \rightarrow 2H^+(aq) + 2e^-$; b 0.0298 mol; c 0.0298 g. **3** a 0.00787 mol; b 1520 C; c 1270 s (21.1 min). **4** a 7200 C; b 0.0746 mol; c 0.925 L. **5** a 0.00750 mol; b 723 C; c 0.301 A. **6** a 0.250 A, 5400 s; b 1350 C; c 1.51 g. **7** 21 600 C; 0.224 mol. **8** 0.560 g. **9** 0.308 L. **10** 0.994 A. **11** 89.4 min. **12** 13.4 h. **13** 16.1 A. **14** 0.771 L. **15** anode $H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$; cathode $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$; overall $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$. **16** 17.2 A. **17** copper needs 2 mol e^- per mole deposited versus 1 mol e^- for silver, so $Q = n(e^-)F$ is twice as large. **18** 2.46 h. **19** 1.61 g. **20** a 0.672 g; b 4.16 L. **21** a 38 600 C; b 71.5 h; c e.g. all of the hydrogen is oxidised (100% fuel utilisation) at constant current, with no side reactions.

Fully-worked solutions

Question 1 a. $t = 20.0\,\text{min} = 1200\,\text{s}$; $Q = It = 3.00 \times 1200 = 3600\,\text{C}$. b. $n(e^-) = \frac{Q}{F} = \frac{3600}{96\,500} = 0.0373\,\text{mol}$. c. From $M(s) \rightarrow M^{2+}(aq) + 2e^-$, $n(M) = \frac{n(e^-)}{2} = \frac{0.037306}{2} = 0.0187\,\text{mol}$.

Question 2 a. $H_2(g) \rightarrow 2H^+(aq) + 2e^-$. b. $t = 12.0\,\text{min} = 720\,\text{s}$; $Q = 4.00 \times 720 = 2880\,\text{C}$;
$n(e^-) = \frac{2880}{96\,500} = 0.0298\,\text{mol}$. c. $n(H_2) = \frac{n(e^-)}{2} = \frac{0.029845}{2} = 0.014922\,\text{mol}$;
$m = n \times M = 0.014922 \times 2.0 = 0.0298\,\text{g}$.

Question 3 a. $n(Cu) = \frac{m}{M} = \frac{0.500}{63.5} = 0.00787\,\text{mol}$. b. From $Cu^{2+} + 2e^- \rightarrow Cu$,
$n(e^-) = 2 \times 0.0078740 = 0.015748\,\text{mol}$; $Q = n(e^-) \times F = 0.015748 \times 96\,500 = 1520\,\text{C}$.
c. $t = \frac{Q}{I} = \frac{1519.7}{1.20} = 1270\,\text{s} = \frac{1266.4}{60} = 21.1\,\text{min}$.

Question 4 a. $t = 20.0\,\text{min} = 1200\,\text{s}$; $Q = 6.00 \times 1200 = 7200\,\text{C}$. b. $n(e^-) = \frac{7200}{96\,500} = 0.0746\,\text{mol}$. c. Anode $H_2 \rightarrow 2H^+ + 2e^-$: $n(H_2) = \frac{0.074611}{2} = 0.037306\,\text{mol}$; $V = n \times V_m = 0.037306 \times 24.8 = 0.925\,\text{L}$.

Question 5 a. $n(Ni) = \frac{0.220}{58.7} = 0.0037479\,\text{mol}$; $n(e^-) = 2 \times 0.0037479 = 0.00750\,\text{mol}$. b.
$Q = n(e^-) \times F = 0.0074957 \times 96\,500 = 723\,\text{C}$. c. $t = 40.0\,\text{min} = 2400\,\text{s}$; $I = \frac{Q}{t} = \frac{723.34}{2400} = 0.301\,\text{A}$.

Question 6 a. $I = 250\,\text{mA} = 0.250\,\text{A}$; $t = 1.50\,\text{h} = 1.50 \times 3600 = 5400\,\text{s}$. b. $Q = It = 0.250 \times 5400 = 1350\,\text{C}$. c. From $Ag^+ + e^- \rightarrow Ag$: $n(e^-) = \frac{1350}{96\,500} = 0.013990\,\text{mol}$, so $n(Ag) = 0.013990\,\text{mol}$; $m = 0.013990 \times 107.9 = 1.51\,\text{g}$.

Question 7 $t = 45.0\,\text{min} = 2700\,\text{s}$; $Q = 8.00 \times 2700 = 21\,600\,\text{C}$; $n(e^-) = \frac{21\,600}{96\,500} = 0.224\,\text{mol}$.

Question 8 $t = 50.0\,\text{min} = 3000\,\text{s}$; $Q = 2.00 \times 3000 = 6000\,\text{C}$; $n(e^-) = \frac{6000}{96\,500} = 0.062176\,\text{mol}$. From $Al \rightarrow Al^{3+} + 3e^-$, $n(Al) = \frac{0.062176}{3} = 0.020725\,\text{mol}$; $m = 0.020725 \times 27.0 = 0.560\,\text{g}$.

Question 9 $t = 25.0\,\text{min} = 1500\,\text{s}$; $Q = 3.20 \times 1500 = 4800\,\text{C}$; $n(e^-) = \frac{4800}{96\,500} = 0.049741\,\text{mol}$. Cathode $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$: $n(O_2) = \frac{0.049741}{4} = 0.012435\,\text{mol}$; $V = 0.012435 \times 24.8 = 0.308\,\text{L}$.

Question 10 $n(\text{Ag}) = \dfrac{2.00}{107.9} = 0.018536\,\text{mol}$; since $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$ is 1:1, $n(\text{e}^-) = 0.018536\,\text{mol}$;

$Q = 0.018536 \times 96\,500 = 1788.7\,\text{C}$; $t = 30.0\,\text{min} = 1800\,\text{s}$; $I = \dfrac{1788.7}{1800} = 0.994\,\text{A}$.

Question 11 $n(\text{H}_2) = \dfrac{0.100}{2.0} = 0.0500\,\text{mol}$; from $\text{H}_2 \rightarrow 2\,\text{H}^+ + 2\,\text{e}^-$, $n(\text{e}^-) = 2 \times 0.0500 = 0.100\,\text{mol}$;

$Q = 0.100 \times 96\,500 = 9650\,\text{C}$; $t = \dfrac{Q}{I} = \dfrac{9650}{1.80} = 5361\,\text{s} = \dfrac{5361}{60} = 89.4\,\text{min}$.

Question 12 $n(\text{H}_2) = \dfrac{5.00}{2.0} = 2.50\,\text{mol}$; $n(\text{e}^-) = 2 \times 2.50 = 5.00\,\text{mol}$; $Q = 5.00 \times 96\,500 = 482\,500\,\text{C}$;

$t = \dfrac{Q}{I} = \dfrac{482\,500}{10.0} = 48\,250\,\text{s} = \dfrac{48\,250}{3600} = 13.4\,\text{h}$.

Question 13 In each minute: $n(\text{H}_2) = \dfrac{0.0100}{2.0} = 0.00500\,\text{mol}$, so $n(\text{e}^-) = 2 \times 0.00500 = 0.0100\,\text{mol}$ and

$Q = 0.0100 \times 96\,500 = 965\,\text{C}$ every 60 s. $I = \dfrac{Q}{t} = \dfrac{965}{60} = 16.1\,\text{A}$.

Question 14 $t = 1.00\,\text{h} = 3600\,\text{s}$; $Q = 5.00 \times 3600 = 18\,000\,\text{C}$; $n(\text{e}^-) = \dfrac{18\,000}{96\,500} = 0.18653\,\text{mol}$. The anode half-equation gives 1 mol CO_2 per 6 mol e^-: $n(\text{CO}_2) = \dfrac{0.18653}{6} = 0.031088\,\text{mol}$; $V = 0.031088 \times 24.8 = 0.771\,\text{L}$.

Question 15 Anode (oxidation): $\text{H}_2(\text{g}) + 2\,\text{OH}^-(\text{aq}) \rightarrow 2\,\text{H}_2\text{O}(\text{l}) + 2\,\text{e}^-$. Cathode (reduction): $\text{O}_2(\text{g}) + 2\,\text{H}_2\text{O}(\text{l}) + 4\,\text{e}^- \rightarrow 4\,\text{OH}^-(\text{aq})$. To combine, the anode half-equation is doubled so that 4 mol e^- appear on each side and cancel, giving $2\,\text{H}_2(\text{g}) + 4\,\text{OH}^-(\text{aq}) \rightarrow 4\,\text{H}_2\text{O}(\text{l}) + 4\,\text{e}^-$. Adding this to the cathode half-equation, the 4 OH^- cancel completely, while the 2 H_2O consumed at the cathode cancel against only 2 of the 4 H_2O formed at the anode — the remaining 2 H_2O are the product:

$$2\,\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\,\text{H}_2\text{O}(\text{l})$$

Question 16 $n(\text{O}_2) = \dfrac{V}{V_m} = \dfrac{0.992}{24.8} = 0.0400\,\text{mol}$; cathode $\text{O}_2 + 4\,\text{H}^+ + 4\,\text{e}^- \rightarrow 2\,\text{H}_2\text{O}$, so

$n(\text{e}^-) = 4 \times 0.0400 = 0.160\,\text{mol}$; $Q = 0.160 \times 96\,500 = 15\,440\,\text{C}$; $t = 15.0\,\text{min} = 900\,\text{s}$; $I = \dfrac{15\,440}{900} = 17.2\,\text{A}$.

Question 17 The charge needed is $Q = n(\text{e}^-) \times F$, and the electrode half-equation fixes how many moles of electrons are required per mole of metal deposited. Silver is deposited by $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$ (1 mol e^- per mol Ag), whereas copper is deposited by $\text{Cu}^{2+} + 2\,\text{e}^- \rightarrow \text{Cu}$ (2 mol e^- per mol Cu). For the same amount, in mole, of metal, copper therefore requires twice as many moles of electrons; because $Q = n(\text{e}^-) \times F$, it requires twice the charge.

Question 18 $n(\text{Zn}) = \dfrac{1.50}{65.4} = 0.022936\,\text{mol}$; $n(\text{e}^-) = 2 \times 0.022936 = 0.045872\,\text{mol}$;

$Q = 0.045872 \times 96\,500 = 4426.6\,\text{C}$; $I = 500\,\text{mA} = 0.500\,\text{A}$; $t = \dfrac{Q}{I} = \dfrac{4426.6}{0.500} = 8853\,\text{s} = \dfrac{8853}{3600} = 2.46\,\text{h}$.

Question 19 $t = 2.00\,\text{h} = 7200\,\text{s}$; $Q = 2.40 \times 7200 = 17\,280\,\text{C}$; $n(\text{e}^-) = \dfrac{17\,280}{96\,500} = 0.17907\,\text{mol}$. From $\text{O}_2 + 4\,\text{H}^+ + 4\,\text{e}^- \rightarrow 2\,\text{H}_2\text{O}$, 4 mol e^- produce 2 mol water, so $n(\text{H}_2\text{O}) = \dfrac{n(\text{e}^-)}{2} = 0.089534\,\text{mol}$;

$m = 0.089534 \times 18.0 = 1.61\,\text{g}$.

Question 20 $t = 90.0\,\text{min} = 5400\,\text{s}$; $Q = 12.0 \times 5400 = 64\,800\,\text{C}$; $n(\text{e}^-) = \dfrac{64\,800}{96\,500} = 0.67150\,\text{mol}$. a. $n(\text{H}_2) = \dfrac{0.67150}{2} = 0.33575\,\text{mol}$; $m = 0.33575 \times 2.0 = 0.672\,\text{g}$. b. $n(\text{O}_2) = \dfrac{0.67150}{4} = 0.16788\,\text{mol}$; $V = 0.16788 \times 24.8 = 4.16\,\text{L}$.

Question 21 a. $n(\text{H}_2) = \dfrac{0.400}{2.0} = 0.200\,\text{mol}$; $n(\text{e}^-) = 2 \times 0.200 = 0.400\,\text{mol}$; $Q = n(\text{e}^-) \times F = 0.400 \times 96\,500 = 38\,600\,\text{C}$. b. $I = 150\,\text{mA} = 0.150\,\text{A}$; $t = \dfrac{Q}{I} = \dfrac{38\,600}{0.150} = 257\,333\,\text{s} = \dfrac{257\,333}{3600} = 71.5\,\text{h}$. c. The calculation assumes all of the hydrogen is oxidised at the anode (100% fuel utilisation) and that the current stays constant at 0.150 A, with no self-discharge or side reactions consuming the hydrogen.

Topic 1 Test A — Energy (multiple choice)

Time allowed: 30 minutes · 25 marks · One scientific calculator permitted · The Chemistry Data Sheet (back of the Practice Examinations booklet) may be used.

Choose the response that is correct or that best answers the question. A correct answer scores 1 mark; an incorrect answer scores 0. Marks are not deducted for incorrect answers. No marks are given if more than one answer is completed for any question. Where required, use $N_A = 6.02 \times 10^{23}\ \text{mol}^{-1}$, $F = 96500\ \text{C mol}^{-1}$, $V_m = 24.8\ \text{L mol}^{-1}$ *at SLC and* $c = 4.18\ \text{J g}^{-1}\,^{\circ}\text{C}^{-1}$.

Selected standard reduction potentials E° *(volts, at SLC):* Cl_2/Cl^- *+1.36;* Ag^+/Ag *+0.80;* $\text{Fe}^{3+}/\text{Fe}^{2+}$ *+0.77;* Cu^{2+}/Cu *+0.34;* Pb^{2+}/Pb *−0.13;* Ni^{2+}/Ni *−0.24;* Cd^{2+}/Cd *−0.40;* Fe^{2+}/Fe *−0.44;* Zn^{2+}/Zn *−0.76;* Al^{3+}/Al *−1.68;* Mg^{2+}/Mg *−2.36.*

Question 1

Which one of the following statements about fuels is correct?

A. Biogas is a fossil fuel because it contains methane.

B. Bioethanol is non-renewable because its combustion releases carbon dioxide.

C. Black coal is a non-renewable fossil fuel that formed over millions of years.

D. Natural gas is a renewable biofuel because it can be supplied continuously by pipeline.

Question 2

Which one of the following correctly describes the energy changes in photosynthesis and cellular respiration?

A. Photosynthesis is endothermic — light energy is stored as chemical energy — while respiration is exothermic, releasing that chemical energy.

B. Both photosynthesis and respiration are exothermic.

C. Photosynthesis is exothermic and respiration is endothermic.

D. Both are endothermic because glucose is produced during photosynthesis.

Question 3

In an endothermic reaction,

A. more energy is released forming the product bonds than is absorbed breaking the reactant bonds.

B. the products store less chemical energy than the reactants.

C. no bonds are broken; energy is only released as bonds form.

D. more energy is absorbed breaking the reactant bonds than is released forming the product bonds, so $\Delta H > 0$.

Question 4

A $3.20\ \text{g}$ sample of a biofuel releases $128\ \text{kJ}$ of heat on complete combustion. Which one of the following is closest to the energy content of the fuel?

A. $40.0\ \text{kJ g}^{-1}$

B. $410\ \text{kJ g}^{-1}$

C. $0.0250\ \text{kJ g}^{-1}$

D. $128\ \text{kJ g}^{-1}$

Question 5

For the reaction $H_2(g) + Cl_2(g) \rightarrow 2\,HCl(g)$, use the bond energies H–H 436, Cl–Cl 242 and H–Cl 431 $kJ\ mol^{-1}$. Which one of the following is closest to ΔH for the reaction?

A. $+184\ kJ\ mol^{-1}$

B. $-184\ kJ\ mol^{-1}$

C. $+247\ kJ\ mol^{-1}$

D. $-862\ kJ\ mol^{-1}$

Question 6

Ethene undergoes complete combustion with $\Delta H = -1411\ kJ\ mol^{-1}$. Which one of the following is closest to the heat released when $0.150\ mol$ of ethene is burned completely?

A. $1411\ kJ$

B. $-212\ kJ$

C. $212\ kJ$

D. $9410\ kJ$

Question 7

Compared with complete combustion, the incomplete combustion of a hydrocarbon fuel

A. releases more energy per mole of fuel.

B. produces only carbon dioxide and water.

C. requires more oxygen per mole of fuel.

D. produces carbon monoxide and/or carbon (soot) and releases less energy per mole of fuel.

Question 8

In a solution calorimeter, $200\ g$ of water is warmed from $20.0\ °C$ to $25.0\ °C$. Which one of the following is closest to the heat energy absorbed by the water?

A. $4.18\ kJ$

B. $20.9\ kJ$

C. $4180\ kJ$

D. $0.0209\ kJ$

Question 9

During the electrical calibration of a calorimeter, $5.00\ kJ$ of energy raises the temperature by $4.00\ °C$. Which one of the following is closest to the calibration factor of the calorimeter?

A. $0.800\ kJ\ °C^{-1}$

B. $1.25\ kJ\ °C^{-1}$

C. $20.0\ kJ\ °C^{-1}$

D. $5.00\ kJ\ °C^{-1}$

Question 10

Using the calorimeter of Question 9 (calibration factor $1.25\,\text{kJ °C}^{-1}$), burning $0.0400\,\text{mol}$ of a fuel raises the temperature by $8.00\,°\text{C}$. Which one of the following is closest to the molar enthalpy of combustion of the fuel?

A. $+250\,\text{kJ mol}^{-1}$

B. $-10.0\,\text{kJ mol}^{-1}$

C. $-250\,\text{kJ mol}^{-1}$

D. $-0.400\,\text{kJ mol}^{-1}$

Question 11

$0.250\,\text{mol}$ of ethanol, C_2H_5OH, undergoes complete combustion. Which one of the following is closest to the volume of carbon dioxide produced at SLC?

A. $6.20\,\text{L}$

B. $24.8\,\text{L}$

C. $18.6\,\text{L}$

D. $12.4\,\text{L}$

Question 12

A snack contains $4.0\,\text{g}$ of fat, $6.0\,\text{g}$ of protein and $25.0\,\text{g}$ of carbohydrate. Using energy values of 37, 17 and $16\,\text{kJ g}^{-1}$ respectively, which one of the following is closest to the energy content of the snack?

A. $560\,\text{kJ}$

B. $650\,\text{kJ}$

C. $1091\,\text{kJ}$

D. $1295\,\text{kJ}$

Question 13

A heater burns fuel that releases $500\,\text{kJ}$ of chemical energy, of which $375\,\text{kJ}$ is transferred to the water being heated. Which one of the following is closest to the energy transformation efficiency?

A. $75.0\,\%$

B. $133\,\%$

C. $25.0\,\%$

D. $0.750\,\%$

Question 14

Which one of the following is the oxidation number of nitrogen in the nitrite ion, NO_2^-?

A. $+5$

B. $+4$

C. $+3$

D. -3

Question 15

In the reaction $\mathrm{Mg(s) + Zn^{2+}(aq) \rightarrow Mg^{2+}(aq) + Zn(s)}$, which species is the oxidant?

A. $\mathrm{Mg(s)}$

B. $\mathrm{Zn(s)}$

C. $\mathrm{Mg^{2+}(aq)}$

D. $\mathrm{Zn^{2+}(aq)}$

Question 16

Which one of the following is the correctly balanced half-equation for the reduction of the iodate ion in acidic solution?

A. $\mathrm{2IO_3^-(aq) + 12H^+(aq) + 5e^- \rightarrow I_2(aq) + 6H_2O(l)}$

B. $\mathrm{2IO_3^-(aq) + 12H^+(aq) + 10e^- \rightarrow I_2(aq) + 6H_2O(l)}$

C. $\mathrm{2IO_3^-(aq) + 6H^+(aq) + 10e^- \rightarrow I_2(aq) + 3H_2O(l)}$

D. $\mathrm{IO_3^-(aq) + 12H^+(aq) + 10e^- \rightarrow I_2(aq) + 6H_2O(l)}$

Question 17

Which one of the following statements about an operating galvanic cell is correct?

A. Oxidation occurs at the anode, which is the negative electrode, and electrons travel through the external wire from the anode to the cathode.

B. Reduction occurs at the anode, which is the positive electrode.

C. Electrons flow through the salt bridge from the anode to the cathode.

D. In the salt bridge, cations migrate toward the anode.

Question 18

A galvanic cell is constructed from a $\mathrm{Cd^{2+}/Cd}$ half-cell and a $\mathrm{Ni^{2+}/Ni}$ half-cell under standard conditions. Which one of the following is closest to the standard cell voltage?

A. $-0.16\ \mathrm{V}$

B. $+0.64\ \mathrm{V}$

C. $+0.16\ \mathrm{V}$

D. $-0.64\ \mathrm{V}$

Question 19

Which one of the following combinations will react spontaneously under standard conditions?

A. $\mathrm{Ag(s)}$ placed in $\mathrm{Zn^{2+}(aq)}$

B. $\mathrm{Zn(s)}$ placed in $\mathrm{Ni^{2+}(aq)}$

C. $\mathrm{Cu(s)}$ placed in $\mathrm{Fe^{2+}(aq)}$

D. $\mathrm{Pb(s)}$ placed in $\mathrm{Mg^{2+}(aq)}$

Question 20

Consider the following standard reduction half-equations and their E° values: $Cl_2 + 2e^- \rightarrow 2Cl^-$ $(+1.36\,V)$; $Ag^+ + e^- \rightarrow Ag$ $(+0.80\,V)$; $Fe^{2+} + 2e^- \rightarrow Fe$ $(-0.44\,V)$; $Al^{3+} + 3e^- \rightarrow Al$ $(-1.68\,V)$. Which species is the strongest oxidant?

A. $Al^{3+}(aq)$

B. $Cl^-(aq)$

C. $Ag^+(aq)$

D. $Cl_2(g)$

Question 21

Two half-cells, Fe^{3+}/Fe^{2+} $(E^\circ = +0.77\,V)$ and Sn^{4+}/Sn^{2+} $(E^\circ = +0.15\,V)$, are connected to form a galvanic cell. Which one of the following is the spontaneous overall cell reaction?

A. $2Fe^{3+}(aq) + Sn^{2+}(aq) \rightarrow 2Fe^{2+}(aq) + Sn^{4+}(aq)$

B. $2Fe^{2+}(aq) + Sn^{4+}(aq) \rightarrow 2Fe^{3+}(aq) + Sn^{2+}(aq)$

C. $Fe^{3+}(aq) + Sn^{2+}(aq) \rightarrow Fe^{2+}(aq) + Sn^{4+}(aq)$

D. $Fe^{2+}(aq) + Sn^{2+}(aq) \rightarrow Fe^{3+}(aq) + Sn^{4+}(aq)$

Question 22

Which one of the following statements about a hydrogen–oxygen fuel cell operating in an acidic electrolyte is correct?

A. Hydrogen is reduced at the cathode.

B. Hydrogen is oxidised at the anode (the negative electrode) and oxygen is reduced at the cathode (the positive electrode).

C. Oxygen is oxidised at the anode.

D. The overall product of the cell is hydrogen gas.

Question 23

Porous electrodes are used in fuel cells mainly to

A. store the gaseous reactants under high pressure.

B. shift the position of equilibrium to favour the product.

C. provide a large surface area for contact between the gas, the electrolyte and the electrode, increasing the reaction rate and efficiency.

D. prevent ions from moving through the electrolyte.

Question 24

A hydrogen–oxygen fuel cell delivers a steady current of $2.50\ \text{A}$ for $15.0\ \text{minutes}$. Which one of the following is closest to the amount, in mol, of electrons that passes through the cell?

A. $3.89 \times 10^{-4}\ \text{mol}$

B. $2250\ \text{mol}$

C. $0.0117\ \text{mol}$

D. $0.0233\ \text{mol}$

Question 25

In a hydrogen–oxygen fuel cell, a steady current of $3.00\ \text{A}$ flows for $1.00\ \text{hour}$. The anode reaction is $\text{H}_2(\text{g}) \rightarrow 2\,\text{H}^+(\text{aq}) + 2\,\text{e}^-$. Which one of the following is closest to the mass of hydrogen consumed at the anode? ($M(\text{H}_2) = 2.0$)

A. $0.224\ \text{g}$

B. $0.112\ \text{g}$

C. $0.0560\ \text{g}$

D. $0.00187\ \text{g}$

End of Topic 1 Test A

Topic 1 Test A — solutions

Question	Answer	Question	Answer
1	C	14	C
2	A	15	D
3	D	16	B
4	A	17	A
5	B	18	C
6	C	19	B
7	D	20	D
8	A	21	A
9	B	22	B
10	C	23	C
11	D	24	D
12	B	25	B
13	A		

Question 1 — C. Coal is a fossil fuel laid down over geological time and not replaced on a human timescale. **A** confuses composition (methane) with origin — biogas is generated from recently living biomass, so it is a renewable biofuel; **B** confuses emitting CO_2 with renewability (bioethanol's crop feedstock regrows); **D** confuses continuous supply through a pipe with renewability.

Question 2 — A. Photosynthesis absorbs light energy and stores it as chemical energy in glucose ($\Delta H > 0$); respiration is the reverse and releases it ($\Delta H < 0$). **B** and **D** wrongly make both the same sign; **C** reverses the two.

Question 3 — D. Endothermic means the energy absorbed to break reactant bonds exceeds the energy released forming product bonds, so $\Delta H > 0$. **A** and **B** describe an exothermic reaction; **C** ignores that bonds must first be broken.

Question 4 — A. Energy content $= \dfrac{128}{3.20} = 40.0\ \text{kJ g}^{-1}$. **B** multiplies ($128 \times 3.20$); **C** inverts the ratio ($3.20 / 128$); **D** forgets to divide by the mass.

Question 5 — B.
$\Delta H = \Sigma(\text{bonds broken}) - \Sigma(\text{bonds formed}) = (436 + 242) - 2(431) = 678 - 862 = -184\ \text{kJ mol}^{-1}$. **A** subtracts in the wrong order (formed − broken); **C** counts only one H–Cl bond ($678 - 431$); **D** counts only the bonds formed.

Question 6 — C. Energy released $= n \times |\Delta H| = 0.150 \times 1411 = 212\ \text{kJ}$ (a positive quantity of heat). **B** wrongly carries the negative ΔH sign into a heat-released magnitude; **A** ignores the mole amount; **D** divides by 0.150 instead of multiplying.

Question 7 — D. Incomplete combustion oxidises carbon only as far as CO or soot, so less energy is released per mole and less oxygen is needed. **A** is the reverse; **B** describes complete combustion; **C** is false (incomplete combustion uses *less* oxygen).

Question 8 — A. $q = m\,c\,\Delta T = 200 \times 4.18 \times (25.0 - 20.0) = 4180\ \text{J} = 4.18\ \text{kJ}$. **B** uses the final temperature ($25.0\,°\text{C}$) instead of $\Delta T = 5.0\,°\text{C}$; **C** reports the joules value as kilojoules; **D** omits the mass.

Question 9 — B. $\text{CF} = \dfrac{E}{\Delta T} = \dfrac{5.00}{4.00} = 1.25\ \text{kJ °C}^{-1}$. **A** inverts the ratio; **C** multiplies $E \times \Delta T$; **D** ignores ΔT.

Question 10 — C. $q = \text{CF} \times \Delta T = 1.25 \times 8.00 = 10.0\,\text{kJ}$; $\Delta H = -\dfrac{q}{n} = -\dfrac{10.0}{0.0400} = -250\,\text{kJ mol}^{-1}$ (exothermic). **A** gives the wrong (positive) sign for a temperature rise; **B** stops at q and forgets to divide by n; **D** multiplies $q \times n$ instead of dividing.

Question 11 — D. $\text{C}_2\text{H}_5\text{OH} + 3\,\text{O}_2 \rightarrow 2\,\text{CO}_2 + 3\,\text{H}_2\text{O}$, so $n(\text{CO}_2) = 2 \times 0.250 = 0.500\,\text{mol}$ and $V = 0.500 \times 24.8 = 12.4\,\text{L}$. **A** uses a $1:1$ ratio; **C** uses the water coefficient $(3:1)$; **B** takes V_m for a single mole.

Question 12 — B. $E = 4.0(37) + 6.0(17) + 25.0(16) = 148 + 102 + 400 = 650\,\text{kJ}$. **C** swaps the fat and carbohydrate energy values; **A** applies $16\,\text{kJ g}^{-1}$ to the whole $35\,\text{g}$; **D** applies $37\,\text{kJ g}^{-1}$ to the whole mass.

Question 13 — A. Efficiency $= \dfrac{\text{useful}}{\text{released}} = \dfrac{375}{500} = 0.750 = 75.0\,\%$. **B** inverts the ratio $(500/375)$; **C** gives the fraction lost; **D** fails to convert the fraction 0.750 into a percentage.

Question 14 — C. In NO_2^-, $x + 2(-2) = -1 \Rightarrow x = +3$. **B** ignores the -1 charge (treating it like neutral NO_2, giving $+4$); **A** is the value in nitrate NO_3^-; **D** has the wrong sign.

Question 15 — D. Zn^{2+} is reduced $(+2 \rightarrow 0)$, so it is the oxidant (it takes electrons). **A** is the reductant (Mg is oxidised); **B** is the reduced product; **C** is the oxidised product.

Question 16 — B. Balancing $2\,\text{IO}_3^- \rightarrow \text{I}_2$: iodine falls $+5 \rightarrow 0$, a gain of $5\,\text{e}^-$ per I, i.e. $10\,\text{e}^-$; $6\,\text{O}$ need $6\,\text{H}_2\text{O}$ and hence $12\,\text{H}^+$. Charge: $-2 + 12 - 10 = 0 = 0$. **A** has too few electrons (5); **C** under-balances H and O; **D** leaves iodine unbalanced (1 vs 2).

Question 17 — A. By definition oxidation is at the anode (the negative terminal of a galvanic cell) and electrons pass through the wire to the cathode. **B** swaps anode and cathode; **C** wrongly sends electrons through the salt bridge (only ions move there); **D** reverses cation migration (cations move to the cathode).

Question 18 — C. The cathode is the higher-E° couple (Ni), the anode is Cd: $E^\circ_{\text{cell}} = E^\circ(\text{cathode}) - E^\circ(\text{anode}) = -0.24 - (-0.40) = +0.16\,\text{V}$. **A** reverses the electrodes; **B** and **D** add the magnitudes $(0.40 + 0.24)$ instead of subtracting.

Question 19 — B. Zn $(E^\circ = -0.76)$ can reduce Ni^{2+} (-0.24): $E^\circ_{\text{cell}} = -0.24 - (-0.76) = +0.52\,\text{V} > 0$, spontaneous. In **A**, **C** and **D** the metal sits below its ion partner in the series, giving $E^\circ_{\text{cell}} < 0$ (Ag cannot reduce Zn^{2+}; Cu cannot reduce Fe^{2+}; Pb cannot reduce Mg^{2+}).

Question 20 — D. The strongest oxidant is the species most readily reduced — the one with the highest E°, namely Cl_2. **A** picks the lowest E° (the weakest oxidant); **B** names the conjugate reductant Cl^-; **C** is a strong but not the strongest oxidant.

Question 21 — A. Fe^{3+} (higher E°) is reduced and Sn^{2+} is oxidised; balancing electrons $(2\,\text{e}^-)$ gives $2\,\text{Fe}^{3+} + \text{Sn}^{2+} \rightarrow 2\,\text{Fe}^{2+} + \text{Sn}^{4+}$, with $E^\circ_{\text{cell}} = +0.62\,\text{V} > 0$. **B** is the non-spontaneous reverse; **C** does not balance electrons $(1\,\text{e}^-$ vs $2\,\text{e}^-)$; **D** is not a balanced redox change.

Question 22 — B. In the acidic H_2–O_2 cell, H_2 is oxidised at the anode $(-)$ and O_2 is reduced at the cathode $(+)$; the product is water. **A** and **C** swap the electrode processes; **D** names the wrong product.

Question 23 — C. Porous electrodes maximise the three-phase contact area between gas, electrolyte and electrode, speeding the electrode reactions. **A** misstates their purpose; **B** confuses them with a catalyst acting on an equilibrium; **D** is the opposite of their role.

Question 24 — D. $Q = It = 2.50 \times (15.0 \times 60) = 2250\,\text{C}$; $n(\text{e}^-) = \dfrac{Q}{F} = \dfrac{2250}{96500} = 0.0233\,\text{mol}$. **A** leaves the time in minutes; **B** forgets to divide by F; **C** divides by $2F$ (a premature stoichiometric factor).

Question 25 — B. $Q = 3.00 \times 3600 = 10800\,\text{C}$; $n(\text{e}^-) = \dfrac{10800}{96500} = 0.1119\,\text{mol}$; $n(\text{H}_2) = 1/2\,n(\text{e}^-) = 0.0560\,\text{mol}$; $m = 0.0560 \times 2.0 = 0.112\,\text{g}$. **A** omits the $\div 2$ from the two electrons per H_2; **C** uses $M = 1.0$ instead of 2.0; **D** leaves the time in minutes.

Topic 1 Test B — Energy (short and extended response)

Time allowed: 50 minutes · 40 marks · One scientific calculator permitted · The Chemistry Data Sheet (back of the Practice Examinations booklet) may be used.

Answer all questions in the spaces provided. Give simplified answers to all numerical questions, with an appropriate number of significant figures. Show all working; no marks are given for an incorrect answer unless accompanied by working. Ensure chemical equations are balanced and that formulas include states, e.g. $\mathrm{H_2(g)}$, $\mathrm{NaCl(s)}$. *Where required, use* $F = 96500\ \mathrm{C\ mol^{-1}}$, $V_m = 24.8\ \mathrm{L\ mol^{-1}}$ *at SLC and* $c = 4.18\ \mathrm{J\ g^{-1}\,°C^{-1}}$.

Question 1 (6 marks)

A student burns ethanol, $\mathrm{C_2H_5OH}$, in a spirit burner to heat $500\ \mathrm{g}$ of water in a metal can. Burning $1.38\ \mathrm{g}$ of ethanol raises the water temperature from $18.0\ °\mathrm{C}$ to $32.0\ °\mathrm{C}$. ($M(\mathrm{C_2H_5OH}) = 46.0$)

a. Write a balanced equation, including states, for the complete combustion of ethanol. 1 mark

b. Calculate the heat energy absorbed by the water. Give your answer in kJ. 2 marks

c. Calculate the experimental molar enthalpy of combustion of ethanol. Give your answer in kJ mol⁻¹. 2 marks

d. The accepted value is $-1367\ \mathrm{kJ\ mol^{-1}}$. Explain why the experimental value is smaller in magnitude than the accepted value. 1 mark

Question 2 (6 marks)

Ethene, $\mathrm{C_2H_4}$, is burned in oxygen. A mixture of $9.00\ \mathrm{L}$ of ethene and $18.0\ \mathrm{L}$ of oxygen, both measured at SLC, is ignited.

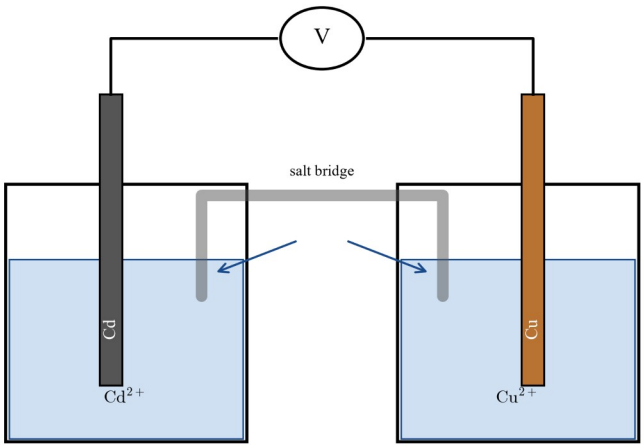
a. Write a balanced equation, including states, for the complete combustion of ethene. 2 marks

b. Determine which reactant is in excess, and by what volume. 2 marks

c. Calculate the volume of carbon dioxide, measured at SLC, produced by the reaction. Give your answer in L. 2 marks

Question 3 (6 marks)

A galvanic cell is set up with a cadmium electrode in $\mathrm{Cd^{2+}(aq)}$ connected to a copper electrode in $\mathrm{Cu^{2+}(aq)}$, as shown. Use $E^\circ(\mathrm{Cd^{2+}/Cd}) = -0.40\ \mathrm{V}$ and $E^\circ(\mathrm{Cu^{2+}/Cu}) = +0.34\ \mathrm{V}$.



a. Using oxidation numbers, identify the species that is oxidised and the oxidant in this cell. 2 marks

b. Write the oxidation and reduction half-equations, including states. 2 marks

c. Calculate the standard cell voltage and state the polarity (sign) of the copper electrode. 2 marks

Question 4 (5 marks)

A hydrogen–oxygen fuel cell supplies a portable device with a steady current of $0.500\ \text{A}$ for 1.80 hours. The anode reaction is $\text{H}_2(\text{g}) \rightarrow 2\,\text{H}^+(\text{aq}) + 2\,\text{e}^-$ and the cathode reaction is $\text{O}_2(\text{g}) + 4\,\text{H}^+(\text{aq}) + 4\,\text{e}^- \rightarrow 2\,\text{H}_2\text{O}(\text{l})$.

a. Calculate the total charge that passes through the cell. Give your answer in C. 1 mark

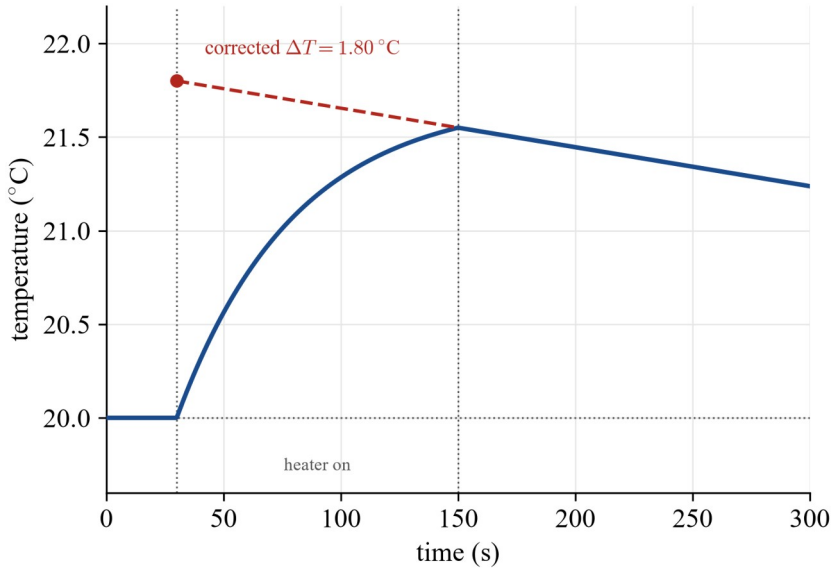
b. Calculate the amount, in mol, of electrons transferred. 1 mark

c. Calculate the mass of hydrogen consumed at the anode. Give your answer in g. ($M(\text{H}_2) = 2.0$) 2 marks

d. Calculate the volume of oxygen, measured at SLC, consumed at the cathode. Give your answer in L. 1 mark

Question 5 (6 marks)

A student calibrates a solution calorimeter by passing an electric current through a heater immersed in the water. The heater draws $2.00\ \text{A}$ at $6.00\ \text{V}$ for $120\ \text{s}$. The temperature–time graph recorded during the calibration is shown.



a. Calculate the electrical energy delivered to the calorimeter ($E = V\,I\,t$). Give your answer in J. 1 mark

b. The corrected temperature rise obtained from the graph is $1.80\,°\text{C}$. Calculate the calibration factor of the calorimeter. Give your answer in J °C^{-1}. 2 marks

c. Another student reads the highest temperature directly from the thermometer instead of extrapolating the cooling line on the graph. State whether this would make the calibration factor too high or too low, and justify your answer. 2 marks

d. The thermometer is graduated at intervals of $0.5\,°\text{C}$. State its resolution and the corresponding reading uncertainty. 1 mark

Question 6 (5 marks)

Glucose, $\text{C}_6\text{H}_{12}\text{O}_6$, is produced in plants by photosynthesis and released as an energy source in cellular respiration, for which $\Delta H = -2803\ \text{kJ mol}^{-1}$. ($M(\text{C}_6\text{H}_{12}\text{O}_6) = 180.0$)

a. Calculate the energy released when $4.50\ \text{g}$ of glucose is completely respired. Give your answer in kJ. 2 marks

b. Photosynthesis is the reverse of respiration. State the sign of ΔH for photosynthesis and explain, in terms of energy, why it has this sign. 1 mark

c. On an energy-profile diagram for the respiration of glucose, state whether the products lie above or below the reactants, and give the value of ΔH that the diagram would show. 2 marks

Question 7 (6 marks)

A city is considering replacing its diesel buses with buses powered by hydrogen–oxygen fuel cells. The hydrogen would be "green" hydrogen, made by electrolysing water using electricity from solar panels.

a. Explain, in terms of energy conversion, why a hydrogen fuel-cell bus can use a greater fraction of its fuel's chemical energy than a diesel engine. 2 marks

b. Explain why hydrogen made by solar-powered electrolysis of water is more sustainable than hydrogen made from natural gas, referring to one green chemistry principle. 2 marks

c. Evaluate the replacement of the diesel buses with fuel-cell buses with respect to sustainability. Refer to two green chemistry principles, state one limitation of the fuel-cell option, and give a concluding recommendation. 2 marks

End of Topic 1 Test B

Topic 1 Test B — solutions

Question 1 $m(\text{water}) = 500\,\text{g}$, $\Delta T = 32.0 - 18.0 = 14.0\,°\text{C}$, $m(\text{ethanol}) = 1.38\,\text{g}$.

(a) $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$. **(1 mark: balanced, with states.)**

(b) $q = mc\Delta T = 500 \times 4.18 \times 14.0 = 29260\,\text{J} = 29.3\,\text{kJ}$. **(1 mark** correct substitution into $q = mc\Delta T$; **1 mark** $29.3\,\text{kJ}$, 3 sf.)

(c) $n(\text{ethanol}) = \dfrac{1.38}{46.0} = 0.0300\,\text{mol}$; the heat released by the fuel is taken as $q = 29.260\,\text{kJ}$, so

$$\Delta H = -\frac{q}{n} = -\frac{29.260}{0.0300} = -975\,\text{kJ mol}^{-1}.$$

$\therefore$ the experimental molar enthalpy of combustion is $-975\,\text{kJ mol}^{-1}$. **(1 mark** $n = 0.0300\,\text{mol}$ and $\Delta H = -q/n$; **1 mark** $-975\,\text{kJ mol}^{-1}$ with the negative sign.)

(d) Heat is lost from the open apparatus to the surroundings (and to the can and escaping hot gases), and combustion may be incomplete, so less than the true amount of heat reaches the water. The measured ΔT, and hence q and $|\Delta H|$, are therefore too small. **(1 mark: heat loss to surroundings / incomplete combustion means the measured value under-estimates the true magnitude.)**

Question 2 Both gases are at SLC, so volume ratios equal mole ratios.

(a) $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$. **(1 mark** correct formulae and balancing; **1 mark** states included.)

(b) To burn $9.00\,\text{L}$ of ethene completely requires $3 \times 9.00 = 27.0\,\text{L}$ of oxygen, but only $18.0\,\text{L}$ is supplied, so **oxygen is the limiting reactant and ethene is in excess**. The oxygen reacts with $\dfrac{18.0}{3} = 6.00\,\text{L}$ of ethene, leaving $9.00 - 6.00 = 3.00\,\text{L}$ of ethene in excess. **(1 mark** identifies oxygen as limiting (ethene in excess) with a supporting comparison; **1 mark** excess ethene $= 3.00\,\text{L}$.)

(c) Carbon dioxide forms from the limiting oxygen in the ratio $O_2 : CO_2 = 3 : 2$:

$$V(CO_2) = 18.0 \times \frac{2}{3} = 12.0\,\text{L at SLC}.$$

(1 mark correct mole/volume ratio from the limiting reactant; **1 mark** $12.0\,\text{L}$.)

Question 3 $E^\circ(Cd^{2+}/Cd) = -0.40\,\text{V}$, $E^\circ(Cu^{2+}/Cu) = +0.34\,\text{V}$.

(a) Cadmium metal has oxidation number 0 and rises to $+2$ in Cd^{2+} — it is **oxidised** (the reductant). Copper falls from $+2$ in Cu^{2+} to 0 in Cu — it is reduced, so Cu^{2+} **is the oxidant**. **(1 mark** Cd oxidised ($0 \rightarrow +2$); **1 mark** oxidant is Cu^{2+} (reduced $+2 \rightarrow 0$).)

(b) Oxidation (anode): $Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$. Reduction (cathode): $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$. **(1 mark** each half-equation, balanced for atoms and charge, with states.)

(c) $E^\circ_{\text{cell}} = E^\circ(\text{cathode}) - E^\circ(\text{anode}) = +0.34 - (-0.40) = +0.74\,\text{V}$. Copper is where reduction occurs, so it is the cathode — the **positive (+)** electrode. **(1 mark** $E^\circ_{\text{cell}} = +0.74\,\text{V}$; **1 mark** copper is positive (the cathode).)

Question 4 $I = 0.500\,\text{A}$, $t = 1.80\,\text{h} = 1.80 \times 3600 = 6480\,\text{s}$.

(a) $Q = It = 0.500 \times 6480 = 3240\,\text{C}$. **(1 mark:** $3240\,\text{C}$**.)**

(b) $n(e^-) = \dfrac{Q}{F} = \dfrac{3240}{96500} = 0.0336\,\text{mol}$. **(1 mark:** $0.0336\,\text{mol}$**, 3 sf.)**

(c) From $H_2 \rightarrow 2H^+ + 2e^-$, $n(H_2) = 1/2\, n(e^-) = 1/2(0.033575) = 0.016788$ mol, so

$$m(H_2) = n \times M = 0.016788 \times 2.0 = 0.0336\,\text{g}.$$

(1 mark $n(H_2) = 1/2\, n(e^-)$; **1 mark** $m = 0.0336$ g.)

(d) From $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$, $n(O_2) = 1/4\, n(e^-) = 0.033575/4 = 0.0083938$ mol, so

$$V(O_2) = n \times V_m = 0.0083938 \times 24.8 = 0.208\,\text{L at SLC}.$$

(1 mark: 0.208 L**.)**

Question 5 $V = 6.00$ V, $I = 2.00$ A, $t = 120$ s.

(a) $E = V\,I\,t = 6.00 \times 2.00 \times 120 = 1440$ J. **(1 mark:** 1440 J**.)**

(b) $\text{CF} = \dfrac{E}{\Delta T} = \dfrac{1440}{1.80} = 800$ J °C^{-1}. **(1 mark** $\text{CF} = E / \Delta T$; **1 mark** 800 J °C^{-1}.)

(c) **Too high.** Because heat is continually lost to the surroundings during the calibration, the highest temperature read directly is lower than the true (heat-loss-corrected) maximum, so the temperature rise used is *too small*. Since $\text{CF} = E / \Delta T$, dividing by a ΔT that is too small gives a calibration factor that is **too high** (e.g. reading the observed rise of about 1.55 °C gives $1440 / 1.55 \approx 929$ J °C^{-1}). **(1 mark** too high; **1 mark** justified — direct reading under-estimates ΔT, and $\text{CF} = E / \Delta T$ rises as ΔT falls.)

(d) The resolution is 0.5 °C (the smallest graduation), which permits readings to ± 0.25 °C (half the smallest division). **(1 mark: resolution** 0.5 °C **with an uncertainty of** ± 0.25 °C**.)**

Question 6 $\Delta H(\text{respiration}) = -2803$ kJ mol^{-1}, $M = 180.0$.

(a) $n(\text{glucose}) = \dfrac{4.50}{180.0} = 0.0250$ mol; energy released $= n \times |\Delta H| = 0.0250 \times 2803 = 70.1$ kJ. **(1 mark** $n = 0.0250$ mol; **1 mark** 70.1 kJ, 3 sf.)

(b) $\Delta H(\text{photosynthesis}) > 0$ (**endothermic**): photosynthesis is the exact reverse of respiration, so it must *absorb* the energy that respiration releases — light energy is taken in and stored as chemical energy in glucose, giving $\Delta H = +2803$ kJ mol^{-1}. **(1 mark:** $\Delta H > 0$**/endothermic because energy (light) is absorbed and stored, the reverse of the exothermic respiration.)**

(c) Respiration is exothermic, so on the energy profile the **products lie below the reactants**, and the diagram shows $\Delta H = -2803$ kJ mol^{-1} (a downward step from reactants to products). **(1 mark** products below reactants; **1 mark** $\Delta H = -2803$ kJ mol^{-1}.)

Question 7

(a) A fuel cell converts the chemical energy of the fuel **directly** into electrical energy in a single controlled redox reaction, so little energy is wasted. A diesel engine converts chemical energy to heat and then to mechanical energy (chemical → heat → mechanical); it is a heat engine whose efficiency is limited, and much of the energy is lost as waste heat. The fuel cell therefore uses a greater fraction of the fuel's chemical energy. **(1 mark** fuel cell = direct chemical → electrical conversion; **1 mark** engine loses energy through the heat step / heat-engine limit.)

(b) Solar-powered electrolysis makes hydrogen from **water using renewable (solar) energy** — a renewable feedstock replenished by natural processes — whereas hydrogen from natural gas relies on a finite fossil resource and releases CO_2 during its production. This illustrates the green chemistry principle of **using renewable feedstocks** (and designing for energy efficiency), so the electrolysis route is more sustainable. **(1 mark** identifies a green chemistry principle (renewable feedstock/energy efficiency); **1 mark** contrasts renewable water/solar with finite CO_2-emitting natural gas.)

(c) *Model answer (rubric).* **Renewable feedstocks:** green hydrogen is generated from water and sunlight, both renewable, replacing finite crude-oil diesel. **Prevention of waste / design for energy efficiency:** the fuel cell emits only water (no CO_2 or particulates) and converts energy more efficiently than a diesel engine. **Limitation:** hydrogen is difficult and costly to store and transport safely (low density, needs high pressure/refuelling infrastructure), and the process is only "green" if the electricity is genuinely renewable. **Conclusion:** provided the hydrogen is produced with renewable electricity, the fuel-cell buses are the more sustainable option and are recommended. **(1 mark** two green chemistry principles applied with evidence; **1 mark** one valid limitation AND a concluding recommendation.)