

Electrochemistry

Mark scheme and worked solutions for 30 questions: 15 multiple choice, 15 structured.

1 Core Electrolysis [1]

Answer: A

Worked solution

Electrolysis always needs three things: an ionic compound, it must be molten or dissolved so the ions can move, and an electric current must be passed through it. Only option A contains all three ideas, so A is correct. The other options describe completely different types of chemical change (thermal decomposition, a metal–acid reaction, and reduction with carbon), none of which involve passing a current through free-moving ions.

2 Core Electrolysis [1]

Answer: B

Worked solution

In any electrolytic cell, there are two electrodes: the anode, which is connected to the positive terminal of the power supply, and the cathode, which is connected to the negative terminal. The question asks for the positive electrode, which is always called the anode, so B is correct.

3 Core Electrolysis [1]

Answer: C

Worked solution

Molten lead(II) bromide, PbBr_2 , contains only Pb^{2+} and Br^- ions — there is no water present. During electrolysis, Pb^{2+} ions move to the cathode and gain electrons to form lead metal (a grey molten bead), while Br^- ions move to the anode and lose electrons to form bromine (a choking, orange-brown vapour/liquid). Since lead is a metal it must appear at the cathode, and bromine, a non-metal, at the anode, which matches option C exactly.

4**Core**

Electrolysis

[1]

Answer: C

Worked solution

Cations (positive ions) are attracted to the cathode; in molten ionic compounds these are always metal ions (or, in some aqueous cases, hydrogen ions), so a metal or hydrogen always forms there. Anions (negative ions) are attracted to the anode; these are always non-metal ions (other than hydrogen), so a non-metal always forms there. This pattern holds for every molten ionic compound, which is exactly what option C states.

5**Core**

Electrolysis

[1]

Answer: A

Worked solution

Dilute sulfuric acid is really a dilute aqueous solution, so the electrolysis is effectively the electrolysis of water: hydrogen collects at the cathode (from reduction of H^+ ions) and oxygen collects at the anode (from oxidation of water/hydroxide ions), in a 2:1 ratio by volume. The sulfate ions stay in solution throughout and become more concentrated as water is used up, but they are not themselves converted into a product, so A is the correct pairing.

6**Core**

Electrolysis

[6]

Mark scheme

(a)

[2]

- cathode (1)
- lead / Pb (1)

(b)

[1]

- bromine / Br₂ (1)

(c)

[2]

- an orange-brown liquid/vapour forms (1)
- a pungent / choking smell is produced (1)

ALLOW fumes seen at the electrode

(d)

[1]

- lead(II) bromide (1)
-

Worked solution

In this cell, lead(II) bromide is molten, so its Pb²⁺ and Br⁻ ions are free to move. The cathode is the negative electrode, and positive Pb²⁺ ions are attracted to it; each Pb²⁺ ion gains two electrons there and is reduced to lead metal, which collects as a silvery liquid bead. The anode is the positive electrode, and negative Br⁻ ions are attracted to it; each pair of Br⁻ ions loses electrons and is oxidised to bromine, seen as an orange-brown liquid or vapour with a sharp, choking smell around the anode. The substance being decomposed throughout — the electrolyte — is the molten lead(II) bromide itself.

7**Core**

Electrolysis

[4]

Mark scheme

(a)

[2]

- decomposition (breakdown) of an ionic compound (1)
- when molten or in aqueous solution, by passing an electric current (1)

(b)(i)

[1]

- cathode (1)

(b)(ii)

[1]

- electrolyte (1)
-

Worked solution

Electrolysis is defined as the decomposition of an ionic compound, when it is molten or dissolved in water, by passing an electric current through it. In this cell, the electrode joined to the negative terminal of the power supply is called the cathode (the positive terminal connects to the anode). The molten lithium chloride itself, which conducts the current and is broken down, is called the electrolyte.

8**Core**

Electrolysis

[1]

Answer: A

Worked solution

A thin layer of chromium on steel cutlery gives it a bright, shiny surface (improving appearance) and forms a protective barrier that stops oxygen and water reaching the steel underneath, which reduces rusting/corrosion. Neither of these purposes relates to weight, cost, melting point, or magnetism, so A is the only correct answer.

9**Core**

Electrolysis

[5]

Mark scheme

(a)

[2]

- improves appearance / makes it look shiny or attractive (1)
- improves resistance to corrosion (1)

(b)

[3]

- the spoon is made the cathode (negative electrode) (1)
- the other electrode (anode) is made of pure silver (1)
- the electrolyte is a solution containing silver ions, e.g. silver nitrate solution (1)

ALLOW any correctly named soluble silver salt solution as the electrolyte

Worked solution

Electroplating coats a cheaper metal object with a thin layer of a more attractive or more corrosion-resistant metal. To electroplate the spoon with silver, the spoon is connected as the cathode, so that positive silver ions in the electrolyte are attracted to it and are reduced (gain electrons) to form a layer of silver metal on its surface. The anode is made of pure silver, which slowly dissolves to replace the silver ions removed from solution. The electrolyte must contain silver ions, so a solution such as silver nitrate is used.

10**Core**

Electrolysis

[1]

Answer: C

Worked solution

Molten calcium bromide, CaBr_2 , contains only Ca^{2+} and Br^- ions. Ca^{2+} ions move to the cathode and are reduced to calcium metal; Br^- ions move to the anode and are oxidised to bromine gas/vapour. Since calcium is the metal and bromine the non-metal, calcium must form at the cathode and bromine at the anode, matching C.

11**Supplement**

Electrolysis

[1]

Answer: B

Worked solution

In aqueous copper(II) sulfate, both Cu^{2+} and H^+ ions are attracted to the cathode, but copper is less reactive than hydrogen, so Cu^{2+} ions are discharged in preference. Each Cu^{2+} ion gains two electrons and is reduced to copper metal, which is deposited on the cathode as a pink/brown (copper-coloured) coating. This matches option B.

12**Supplement**

Electrolysis

[7]

Mark scheme

(a) [1]

- copper (deposited on the cathode) (1)

(b) [1]

- oxygen (1)

(c) [2]

- the copper anode dissolves (1)
- copper atoms lose electrons and enter the solution as Cu^{2+} ions (1)

(d) [3]

- with copper electrodes: Cu^{2+} ions are removed from solution at the cathode (as copper is deposited) (1)
- but are replaced at the same rate by Cu^{2+} ions entering solution from the dissolving copper anode, so the concentration of Cu^{2+} (and the blue colour) stays roughly constant (1)
- with carbon electrodes: Cu^{2+} ions are removed from solution at the cathode but are not replaced, because the inert anode does not dissolve (it produces oxygen instead), so the concentration of Cu^{2+} falls and the blue colour fades (1)

Worked solution

In both experiments, Cu^{2+} ions are reduced at the cathode and copper metal is deposited there, so the cathode product is the same either way. The difference is at the anode. With inert carbon electrodes, the anode cannot react, so hydroxide ions/water are oxidised instead, releasing oxygen gas; this uses up Cu^{2+} ions from solution without replacing them, so the blue colour gradually fades. With copper electrodes, the anode itself is oxidised — copper atoms lose electrons and dissolve into solution as new Cu^{2+} ions — replacing the Cu^{2+} ions removed at the cathode at almost the same rate, so the solution's colour and concentration stay roughly constant. This is the basis of the industrial purification of copper.

13**Supplement**

Electrolysis

[1]

Answer: D

Worked solution

At the anode, negative ions lose electrons (oxidation); these electrons flow away from the anode, through the external wires and the battery, and back to the cathode, where positive ions gain them (reduction). So although current directions are sometimes described differently, the electrons themselves move from the anode, through the external circuit, to the cathode — this is option D. Inside the electrolyte, by contrast, charge is carried by the ions moving towards their respective electrodes, not by free electrons.

14

Supplement

Electrolysis

[6]

Mark scheme

(a)

[2]

- $\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$ (2)

ALLOW correct species (Mg^{2+} , e^- , Mg) present but unbalanced (1)

(b)

[2]

- $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$ (2)

ALLOW correct species present but unbalanced (1)

(c)(i)

[1]

- ions move through the electrolyte / cations move to the cathode and anions move to the anode (1)

(c)(ii)

[1]

- electrons flow through the wires, from the anode to the cathode (via the power supply) (1)

Worked solution

At the cathode, each Mg^{2+} ion gains two electrons and is reduced to magnesium metal: $\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$. At the anode, pairs of Cl^- ions lose electrons and are oxidised to chlorine gas: $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$. Charge is carried through the whole circuit in two different ways: inside the molten electrolyte, it is carried by the ions themselves moving towards the oppositely-charged electrode (Mg^{2+} towards the cathode, Cl^- towards the anode); outside the electrolyte, in the wires and battery, charge is carried by electrons, which flow from the anode (where they are released) to the cathode (where they are needed), driven by the power supply.

15

Supplement

Electrolysis

[1]

Answer: C**Worked solution**

When a halide solution is concentrated, there is a very high concentration of halide ions compared with hydroxide ions, so the halide ion is discharged at the anode instead of the usual oxygen. Here, that means Br^- ions lose electrons and are oxidised to bromine, so C is correct. Potassium (option D) can never be discharged in aqueous solution because it is far more reactive than hydrogen.

16**Supplement**

Electrolysis

[7]

Mark scheme

(a) [2]

- hydrogen (1)
- sodium is more reactive than hydrogen, so H^+ ions are discharged (reduced) in preference to Na^+ ions (1)

(b) [1]

- oxygen (1)

(c) [1]

- iodine (1)

(d) [3]

- in the dilute solution, iodide ions are present at a much lower concentration than hydroxide ions/water molecules, so hydroxide ions/water are preferentially discharged, producing oxygen (1)
- in the concentrated solution, iodide ions are present at a much higher concentration, so iodide ions are preferentially discharged instead, producing iodine (1)
- at high enough concentration, the more concentrated ion can be discharged in preference to the ion that is normally favoured (1)

Worked solution

In both experiments, the cathode product is the same: hydrogen, because sodium is far more reactive than hydrogen and so Na^+ ions are never discharged in aqueous solution — H^+ ions are reduced instead. At the anode, the outcome depends on concentration. In the dilute solution, there are relatively few I^- ions compared with the water/ OH^- present, so oxygen is produced from the oxidation of water/hydroxide ions. In the concentrated solution, there are now far more I^- ions available, so these are discharged in preference, producing iodine. This shows that selective discharge depends not only on general reactivity but also on the relative concentration of ions in solution.

17**Supplement**

Electrolysis

[1]

Answer: D**Worked solution**

At the anode, oxidation occurs, meaning ions lose electrons. Chloride ions do this in pairs: $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$, which is balanced for both chlorine atoms and charge (-2 on the left equals -2 from the two electrons released on the right). Option A shows potassium gaining an electron, which is a reduction reaction happening at the cathode, not the anode. Option B reverses the correct anode equation into a reduction. Option C is not balanced in charge, since 2Cl^- (charge -2) plus 2e^- (charge -2) would give a total charge of -4 , which cannot equal the 0 charge of Cl_2 .

18

Supplement

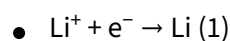
Electrolysis

[6]

Mark scheme

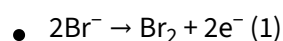
(a)(i)

[1]



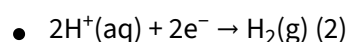
(a)(ii)

[1]



(b)(i)

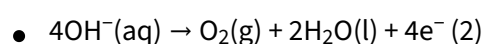
[2]



ALLOW correct species without state symbols (1)

(b)(ii)

[2]



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

ALLOW correct species without state symbols, or unbalanced (1)

Worked solution

For molten lithium bromide, each Li^+ ion gains one electron at the cathode to form lithium metal, and pairs of Br^- ions lose electrons at the anode to form bromine: $\text{Li}^+ + \text{e}^- \rightarrow \text{Li}$ and $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$. For dilute sulfuric acid, hydrogen ions are reduced at the cathode: $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$. At the anode, since sulfate ions are not discharged, oxygen is produced from the oxidation of hydroxide ions or water molecules; this can be written either as $4\text{OH}^-(\text{aq}) \rightarrow \text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^-$, or equivalently starting from water molecules directly: $2\text{H}_2\text{O}(\text{l}) \rightarrow \text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^-$. Both versions are balanced for atoms and for charge.

19

Core

Hydrogen–oxygen fuel cells

[1]

Answer: A**Worked solution**

Inside the fuel cell, hydrogen and oxygen react to form water and release electrical energy directly, without burning any fuel. Because no carbon-containing substance is involved anywhere in this reaction, water is the only chemical product — matching option A. Carbon dioxide only comes from burning carbon-based fuels such as petrol, which is not what happens in a fuel cell.

20**Core**

Hydrogen–oxygen fuel cells

[1]

Answer: B

Worked solution

The name "hydrogen–oxygen fuel cell" tells you exactly which two gases are supplied and react: hydrogen and oxygen. They combine to release electrical energy and form water, matching option B. None of the other options name this correct pair of reactants.

21**Core**

Electrolysis

[1]

Answer: B

Worked solution

Because the sodium chloride solution is concentrated, there is a very high concentration of Cl^- ions present, so these are oxidised at the anode in preference to hydroxide ions/water, releasing chlorine gas. This matches B. At the cathode (not asked here), hydrogen would be produced instead of sodium, since sodium is far too reactive to be discharged from aqueous solution.

22**Core**

Electrolysis

[1]

Answer: D

Worked solution

Molten strontium chloride, SrCl_2 , contains only Sr^{2+} and Cl^- ions. Sr^{2+} ions are reduced at the cathode to form strontium metal, and Cl^- ions are oxidised at the anode to form chlorine gas. Since strontium is the metal, it must appear at the cathode, and chlorine, the non-metal, at the anode — matching D. This is exactly the same pattern as for any other molten binary metal halide.

23**Core**

Electrolysis

[5]

Mark scheme

(a)

[2]

- hydrogen (1)
- hold a lit/burning splint near the gas; it burns/ignites with a squeaky pop (1)

(b)

[2]

- oxygen (1)
- hold a glowing splint in the gas; it relights (1)

(c)

[1]

- dilute sulfuric acid (1)
-

Worked solution

Dilute sulfuric acid behaves like acidified water when electrolysed, so hydrogen is produced at the cathode and oxygen at the anode, in a 2:1 ratio by volume. Hydrogen can be identified by collecting a sample and bringing a lit splint near it: it burns rapidly, giving a characteristic squeaky pop. Oxygen can be identified by inserting a glowing (not flaming) splint into the gas: it relights, bursting back into flame, because oxygen supports combustion. The substance being electrolysed throughout — the electrolyte — is the dilute sulfuric acid itself.

24**Core**

Electrolysis

[5]

Mark scheme

(a)

[2]

- improves appearance / gives a shiny, attractive finish (1)
- improves resistance to corrosion/rusting (1)

(b)

[3]

- the bumper is made the cathode (negative electrode) (1)
 - the other electrode (anode) is made of pure chromium (1)
 - the electrolyte is a solution containing chromium ions, e.g. a chromium salt solution (1)
-

Worked solution

Chromium plating gives steel a bright, shiny surface and protects the steel underneath from oxygen and moisture, reducing corrosion. To carry out the plating, the steel bumper is connected as the cathode, so that positive chromium ions in the electrolyte are attracted to it and reduced to form a layer of chromium metal on its surface. The anode is made of pure chromium, which dissolves to replace the chromium ions used up at the cathode. The electrolyte must therefore contain chromium ions, such as in a chromium salt solution.

Mark scheme

(a) [2]

- $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ (2)

ALLOW correct species present but unbalanced (1)

(b) [2]

- $\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$ (2)

ALLOW correct species present but unbalanced (1)

(c) [2]

- mass of the anode decreases, because copper dissolves from it into solution as Cu^{2+} ions (1)
- mass of the cathode increases, because copper is deposited onto it (1)

(d) [2]

- Cu^{2+} ions are removed from solution at the cathode (1)
- Cu^{2+} ions are replaced at the same rate by Cu^{2+} ions entering solution from the dissolving (impure copper) anode (1)

Worked solution

At the cathode, Cu^{2+} ions from solution are reduced onto the pure copper strip: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$, so the cathode gains mass as it becomes coated with newly deposited copper. At the anode, copper atoms from the impure block are oxidised into solution: $\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$, so the anode loses mass as it dissolves away (leaving behind any impurities that do not dissolve, which fall to the bottom as "anode sludge"). Because Cu^{2+} ions are added to the solution at the anode at essentially the same rate as they are removed at the cathode, the overall concentration of copper(II) sulfate solution stays roughly the same throughout the process. This is the basis of the industrial electrolytic purification (refining) of copper.

Mark scheme

(a) [2]

- cathode: hydrogen (1)
- anode: chlorine (1)

(b) [2]

- $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$ (2)

ALLOW correct species present but unbalanced (1)

(c) [2]

- sodium hydroxide solution (1)
- valid industrial use, e.g. making soap or detergents, or as an industrial alkali, or bleach manufacture (1)

(d) [2]

- the brine is concentrated, so the concentration of Cl^- ions is very high (1)
- at such a high concentration, Cl^- ions are discharged (oxidised) in preference to hydroxide ions/water, overriding the usual preference for oxygen seen in dilute solution (1)

Worked solution

Electrolysing concentrated brine produces hydrogen at the cathode (from reduction of H^+ ions/water, since sodium is too reactive to be discharged) and chlorine at the anode: $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$. Sodium ions and the remaining hydroxide ions stay behind in solution, so the solution left over is sodium hydroxide, which is an important industrial alkali used, for example, in soap and detergent manufacture. The reason chlorine rather than oxygen forms at the anode is concentration: in a dilute solution, oxygen would normally form because hydroxide ions/water are usually easier to discharge than chloride ions, but in this highly concentrated brine there are so many more Cl^- ions available that they are discharged instead, overriding the usual preference.

Mark scheme

(a) [2]

- charge is carried by the movement of ions (1)
- positive ions (H^+) move towards the cathode and negative ions (SO_4^{2-}) move towards the anode (1)

ALLOW OH^- (from water) as the negative ion moving towards the anode

(b) [1]

- charge is carried by the movement/flow of electrons through the wires (1)

(c) [2]

- hydrogen ions gain electrons (1)
- they are reduced, forming hydrogen gas / H_2 (1)

Worked solution

Inside the electrolyte, there are no free electrons — instead, charge is carried by the ions themselves as they move: positive hydrogen ions move towards the negative cathode, and negative sulfate ions move towards the positive anode. Outside the electrolyte, in the wires of the external circuit, there are no free ions — instead, charge is carried by electrons flowing through the metal wire, driven by the battery. When hydrogen ions reach the cathode, they gain the electrons supplied by the circuit and are reduced to hydrogen gas.

Mark scheme

(a)

[1]

- hydrogen and oxygen (1)

(b)

[3]

- water (1)
- the fuel cell reaction does not involve burning/combustion of a carbon-based fuel (1)
- hydrogen combines directly with oxygen to release electrical energy, forming only water, with no carbon present anywhere in the reaction to form carbon dioxide (1)

ALLOW petrol is a hydrocarbon / contains carbon, so burning it produces carbon dioxide, whereas hydrogen contains no carbon (1)

Worked solution

The fuel cell is supplied with hydrogen and oxygen gas. Inside the cell, these combine to produce electricity, with water as the only chemical product — no combustion of a carbon-containing fuel takes place, so no carbon dioxide can be formed at any point. This is different from a petrol engine, which burns a hydrocarbon fuel and always produces carbon dioxide (and often other pollutants) as combustion products, since petrol contains carbon.

Mark scheme

(a) [2]

- any two from: (1 each)
- produces no carbon dioxide / no polluting exhaust gases, only water (1)
- converts chemical energy to useful (electrical) energy more efficiently than a petrol engine, which loses more energy as heat (1)
- quieter running / fewer moving parts (1)
- hydrogen can be made from water (by electrolysis) so is a renewable fuel if renewable electricity is used (1)

ALLOW no dependence on crude oil / petrol, a non-renewable resource

(b) [2]

- any two from: (1 each)
- hydrogen is difficult/expensive to store safely, e.g. needs high-pressure tanks or very low temperatures, and takes up more space than the same energy stored as petrol (1)
- hydrogen does not occur freely in nature and must be manufactured, often using energy from fossil fuels (1)
- hydrogen is flammable and poses an explosion risk (1)
- fuel cells / hydrogen vehicles are expensive to manufacture, e.g. platinum catalyst (1)
- few hydrogen refuelling stations (1)

ALLOW shorter range between refuelling

(c) [2]

- a specific reason, e.g. there are very few hydrogen refuelling stations compared with petrol stations / fuel cell vehicles and hydrogen fuel are currently expensive / safe hydrogen storage technology is still being developed (1)
- linked consequence, e.g. so drivers cannot easily refuel / are unwilling to buy the vehicles (1)

ALLOW a reason already credited in (b) may be used again here**IGNORE** vague answers with no specific reason (0)**Worked solution**

Compared with a petrol engine, a hydrogen fuel cell has real advantages: it produces no carbon dioxide or other polluting gases as it runs, since its only chemical product is water, and it can convert chemical energy into useful energy more efficiently, since a fuel cell does not waste as much energy as heat as a combustion engine does. However, there are real disadvantages too: hydrogen gas is bulky and difficult to store safely, usually needing heavy high-pressure tanks, and it is not found freely in nature, so producing it usually requires energy, which may itself come from fossil fuels. These practical difficulties — combined with the small number of hydrogen refuelling stations currently available and the higher

cost of the technology — help explain why hydrogen fuel cell vehicles are not yet as common as petrol vehicles, even though the fuel cell itself is a clean way of generating electricity.

30 **Supplement** Hydrogen–oxygen fuel cells

[6]

Mark scheme

(a) [2]

- only water (vapour) leaves the exhaust (1)
- no carbon dioxide, carbon monoxide or other polluting gases are released while driving, because no carbon-based fuel is burned (there is no combustion) (1)

(b) [4]

- the claim is only true for emissions "at the tailpipe" while driving, not for the whole life cycle of the fuel (1)
- hydrogen does not occur freely in useful quantities and must be manufactured, e.g. by the electrolysis of water, or from natural gas/hydrocarbons (1)
- if the electricity or process used to manufacture the hydrogen comes from burning fossil fuels, carbon dioxide and other pollutants are released during production (1)
- so the "zero emissions" claim is misleading for the fuel's whole life cycle unless the hydrogen is made using a low-carbon/renewable energy source, even though the car itself is genuinely pollution-free while driving (1)

ALLOW valid alternative reasoning, e.g. energy/emissions involved in manufacturing or disposing of the fuel cell and storage tank

Worked solution

While the car is being driven, the claim is genuinely true: hydrogen and oxygen combine in the fuel cell to form only water, so the exhaust really does release no carbon dioxide or other pollutants — the car itself is pollution-free in use. However, "100% pollution-free" is a much bigger claim than that, because it ignores where the hydrogen fuel comes from. Hydrogen is not a naturally occurring fuel source; it has to be manufactured, commonly by electrolysis of water or by processing natural gas, both of which require energy. If that energy comes from burning fossil fuels, then pollutant gases, including carbon dioxide, are released during hydrogen production, even though none come from the car itself. So the advertisement's claim is fair only for the car's own emissions while driving, not for the environmental impact of the whole process of making its fuel — a more honest claim would depend on how "green" the hydrogen supply actually is.