SMART EXAM RESOURCES

IGCSE CHEMISTRY 0620 /0971 (9-1)

[SPECIAL -MICRO TOPIC QUESTIONS]

ELECTROLYSIS- [free sample copy]

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SMART EXAM RESOURCES

SUBJECT: CHEMISTRY

TOPIC:ELECTROCHEMISTRY

SET-1-QP-MS

ELECTROLYSIS BASIC QUESTIONS

1 State **two** properties of graphite (carbon) which make it suitable for use as an electrode.

1 .....

2 .....

[2]

MARK SCHEME:

	M1 inert (1) **M2** conducts electricity (1)	**2**

2 This question is about electrolysis.

(a) State the meaning of the term *electrolyte*.

.....

..... [2]

MARK SCHEME:

Question	Answer	Marks
(a)	**M1** ionic compound **AND** either molten or aqueous(or both)(1) **M2** conducts electricity / undergoes electrolysis(1)	**2**

3 This question is about electricity and chemical reactions.

(a) Aqueous copper(II) sulfate is an electrolyte.
The electrolysis of aqueous copper(II) sulfate using inert electrodes forms:

- copper at the cathode
- oxygen at the anode.

(i) State what is meant by the term electrolyte.

.....

..... [2]

MARK SCHEME:

Question	Answer	Marks
(a)(i)	**M1** ionic compound **M2** molten and / or aqueous	2

4 Molten potassium chloride undergoes electrolysis.

(i) State what is meant by the term *electrolysis*.

.....

..... [2]

MARK SCHEME:

(i)	**breakdown** by (the passage of) **electricity** (1) of an **ionic compound** in **molten or aqueous** (state) (1)	**2**

5 **a** State **two** reasons why carbon (graphite) is suitable to use as an electrode.

1 .....

2 .....
[2]

b Name the particle responsible for the conduction of electricity in the metal wires used in a circuit.

..... [1]

MARK SCHEME:

Question	Answer	Marks
a	**M1** inert (1) **M2** good conductor of electricity (1)	2
b	electron	1

ELECTROPLATING

5.1.1

One of the methods used to prevent iron or steel from rusting is to electroplate it with another metal, such as tin. Complete the following.

The anode is made of ..... .

The cathode is made of ..... .

The electrolyte is a solution of ..... .

[3]

[3]

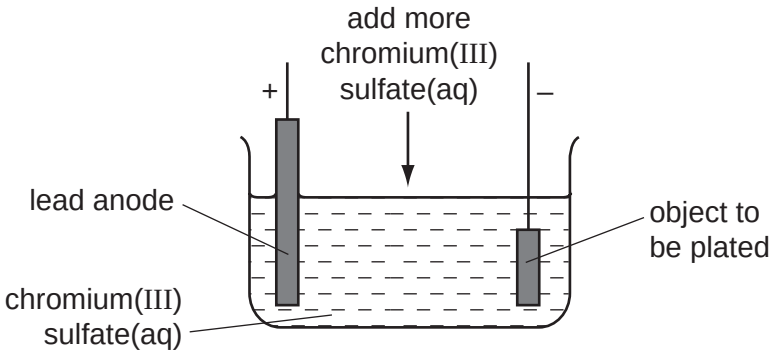
----------Marking Scheme----------

anode	tin	NOT impure time	[1]
cathode	iron or steel		[1]
tin salt **or** tin ions as electrolyte NOT oxide or hydroxide or carbonate			[1]

5.1.2

Chromium is a transition element.

Chromium is used to electroplate steel objects. The diagram shows how this could be done



(i) Give **two** reasons why steel objects are plated with chromium.

.....

..... [2]

(ii) The formula of the chromium(III) ion is Cr^{3+} and of the sulfate ion is SO_4^{2-}. Give the formula of chromium(III) sulfate.

..... [1]

(iii) Write the equation for the reaction at the negative electrode (cathode).

..... [2]

(iv) A colourless gas, which relights a glowing splint, is formed at the positive electrode (anode). Name this gas.

.....

.....

..... [1]

(v) During electrolysis, it is necessary to add more chromium(III) sulfate but during copper-plating using a copper anode, it is not necessary to add more copper(II) sulfate. Explain.

.....

.....

----------Marking Scheme----------

(i) appearance/shiny/more attractive/decoration
resist corrosion / rusting
hard surface
any **TWO** [2]
NOT becomes harder / stronger

(ii) $Cr_2(SO_4)_3$ [1]
ignore correct charges on ions

(iii) $Cr^{3+} + 3e \rightarrow Cr$ [2]
Cr^{3+} to Cr only [1]
ignore comments about sulfate ion

(iv) oxygen / O_2 [1]

(v) to replace chromium ions (used to plate steel) [1]
/ chromium sulfate used up

copper ions replaced from copper anode [1]
/ solution of copper sulfate does not change
not just that anode is not made of chromium

ELECTROPLATING

1 Metal objects can be electroplated with silver.

(i) Describe how a metal spoon can be electroplated with silver.
Include:

- what to use as the positive electrode and as the negative electrode
- what to use as the electrolyte
- an ionic half-equation to show the formation of silver.

You may include a diagram in your answer.

.....

.....

.....

ionic half-equation .....

[4]

(ii) Give **one** reason why metal spoons are electroplated with silver.

.....

..... [1]

MARKING SCHEME:

(i)	**M1** spoon as cathode **M2** (pure)silver as anode **M3** aqueous silver nitrate as electrolyte **M4** $Ag^+ + e^- \rightarrow Ag$	**4**
(ii)	any one from: ∞ Improves appearance ∞ prevent / resist corrosion / oxidation ∞ antibacterial	**max 1**

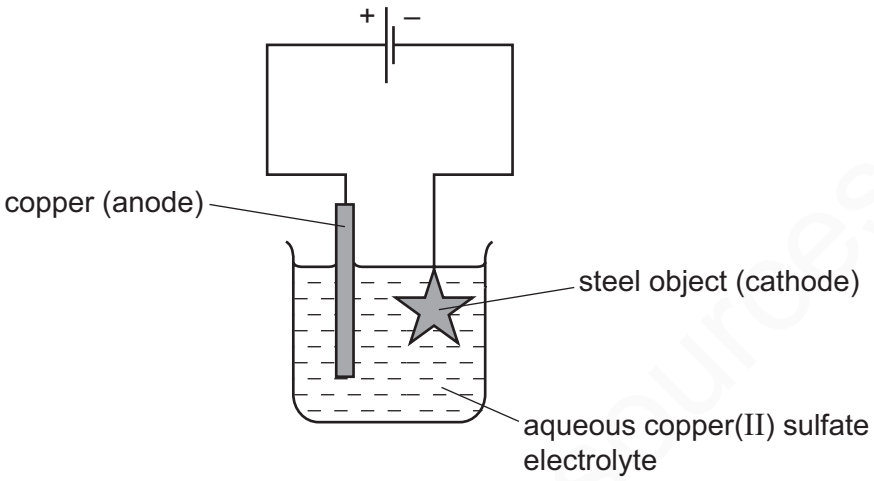
2 Electroplating steel objects with silver involves a three-step process.

step 1 A coating of copper is applied to the object.

step 2 A coating of nickel is applied to the object.

step 3 The coating of silver is applied to the object.

(a) A diagram of the apparatus used for **step 1** is shown.



(i) The chemical process taking place on the surface of the object is

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

Explain whether this process is oxidation or reduction.

.....

..... [1]

(ii) Explain why the concentration of copper ions in the electrolyte remains constant throughout **step 1**.

.....

.....

..... [2]

(b) Give **two** changes which would be needed in order to coat nickel onto the object in **step 2**.

.....

.....

.....

..... [2]

(c) Copper, nickel and silver are transition elements.
Typical physical properties of transition elements are a high density and a high melting point.

Give **three** different properties of transition metals which are not typical of other metals.

.....

.....

..... [3]

[Total: 8]

MARKING SCHEME:

(a)(i)	reduction **and** (the Cu^{2+} ion/copper ions) is gaining electrons/is decreasing in oxidation number;	**1**
(a)(ii)	formation of Cu^{2+}/copper ions at the anode happens at the same rate as; removal of Cu^{2+}/copper ions at the cathode ora;	**2** 1 1
(b)	replace (anode of) copper with nickel; replace electrolyte with nickel(II) sulfate/$NiSO_4$;	**2** 1 1
(c)	(good) catalysts; variable oxidation numbers; form coloured compounds/coloured ions;	**3** 1 1 1

SMART EXAM RESOURCES
SUBJECT: CHEMISTRY
TOPIC:ELECTROCHEMISTRY
SET-3-QP-MS

ELECTROPLATING

1 A metal spoon is electroplated with copper.

State what is used as:

the positive electrode (anode) .....

the negative electrode (cathode) .....

the electrolyte. .....

[3]

MARK SCHEME:

	copper (1) spoon (1) (aqueous or solution) of named copper salt (1)	**3**

2 Spoons can be electroplated with silver.

(i) Name the substances used as:

the anode (positive electrode) .....

the cathode (negative electrode) .....

the electrolyte. .....

[3]

(ii) State **two** reasons why spoons are electroplated.

1 .....

2 .....

[2]

MARK SCHEME:

(i)	**M1** silver **M2** spoon **M3** (aqueous or solution) of silver nitrate	**3**
(ii)	**M1** prevent corrosion **M2** improve appearance	**2**

CATHODIC PROTECTION

5.3.1

(a) Titanium is produced by the reduction of its chloride. This is heated with magnesium in an inert atmosphere of argon.

$$TiCl_4 + 2Mg \rightarrow Ti + 2MgCl_2$$

(i) Explain why it is necessary to use argon rather than air.

..... [1]

(ii) Name another metal that would reduce titanium chloride to titanium.

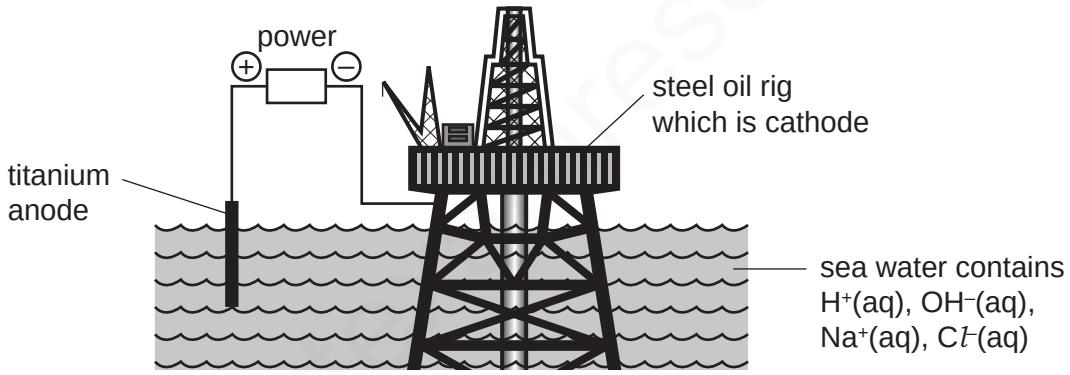
..... [1]

(iii) Suggest how you could separate the metal, titanium, from the soluble salt magnesium chloride.

.....

..... [2]

(b) Titanium is very resistant to corrosion. One of its uses is as an electrode in the cathodic protection of large steel structures from rusting.



(i) Define oxidation in terms of electron transfer.

..... [1]

(ii) The steel oil rig is the cathode. Name the gas formed at this electrode.

..... [1]

(iii) Name the **two** gases formed at the titanium anode.

..... and ..... [2]

(iv) Explain why the oil rig does not rust.

.....

..... [2]

(v) Another way of protecting steel from corrosion is sacrificial protection.
Give **two** differences between sacrificial protection and cathodic protection.

.....

..... [2]

[Total: 12]

(a) (i) air would react (with the magnesium **or** titanium)
OR argon would not react (with the metals) [1]
NOT argon is inert

(ii) any metal higher than magnesium in reactivity series [1]

(iii) add water (to dissolve salt) [1]
filter **or** centrifuge [1]

(b) (i) electron loss [1]

(ii) hydrogen [1]

(iii) oxygen [1]
chlorine [1]

(iv) it cannot lose electrons (because) [1]
it receives electrons (from the battery) [1]

OR reduction occurs at the cathode [1]
oxidation at the anode (not cathode) [1]

OR electrons are "pushed" to rig [1]
preventing it from being oxidised [1]

for comments of the type – rusting needs oxygen, it is formed on titanium not iron **ONLY** [1]

NOT the idea that titanium is more reactive etc

(v) **SET 1**
sacrificial protection is a cell
does not need electricity
cathodic protection is electrolysis
cathodic protection needs electricity

SET 2
sacrificial protection needs a more reactive metal (in contact with iron or steel)
this metal corrodes instead of steel
cathodic protection needs an inert electrode accept unreactive or less reactive metal as an electrode
has to be **ONE** comment from each set [2]
all comments about oxide layers and coating are neutral

[Total: 12]

ELECTROLYSIS OF COPPER SULFATE

5.5.1

Copper has the structure of a typical metal. It has a lattice of positive ions and a "sea" of mobile electrons. The lattice can accommodate ions of a different metal.

(a) Aqueous copper(II) sulphate solution can be electrolysed using carbon electrodes. The ions present in the solution are as follows.

$Cu^{2+}(aq)$, $SO_4^{2-}(aq)$, $H^+(aq)$, $OH^-(aq)$

(i) Write an ionic equation for the reaction at the negative electrode (cathode).

.....

[1]

(ii) A colourless gas was given off at the positive electrode (anode) and the solution changes from blue to colourless.

Explain these observations.

.....

.....

[2]

(b) Aqueous copper(II) sulphate can be electrolysed using copper electrodes. The reaction at the negative electrode is the same but the positive electrode becomes smaller and the solution remains blue.

(i) Write a word equation for the reaction at the positive electrode.

.....

[1]

(ii) Explain why the colour of the solution does not change.

.....

.....

[2]

(iii) What is the large scale use of this electrolysis?

.....

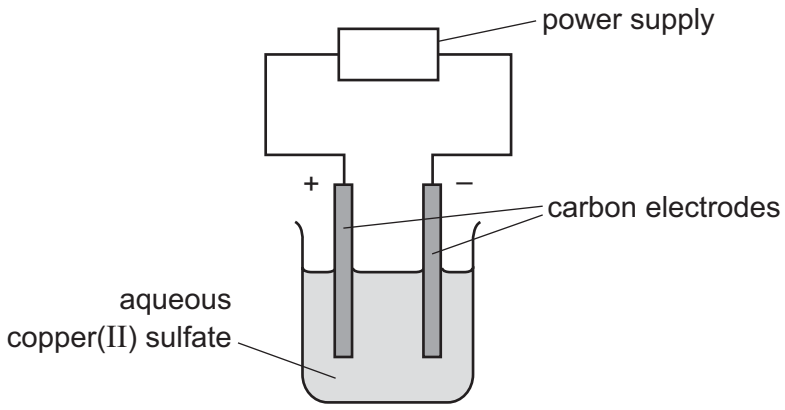
[1]

----------Marking Scheme----------

(a)	(i)	$Cu^{2+} + 2e = Cu$	[1]
	(ii)	gas is oxygen	[1]
		(copper(II) sulphate) changes to <u>sulphuric acid</u> **or** copper ions removed from solution	[1]
(b)	(i)	copper atoms - electrons = copper ions accept correct symbol equation	[1]
	(ii)	concentration of copper ions does not change **or** amount **or** number of copper ions does not change	[1]
		copper ions are removed and then replaced **or** copper is transferred from anode to cathode	[1]
	(iii)	refining copper **or** plating (core) **or** extraction of boulder copper	[1]

ELECTROLYSIS OF COPPER SULFATE

1 A student electrolyses aqueous copper(II) sulfate using the apparatus shown.



Oxygen gas forms at the positive electrode (anode).

(i) Write an ionic half-equation for the reaction at the negative electrode (cathode). Include state symbols.

..... [3]

(ii) Describe what the student observes at the negative electrode.

..... [1]

(iii) Give **two other** observations which the student makes during the electrolysis.

1 .....

2 .....
[2]

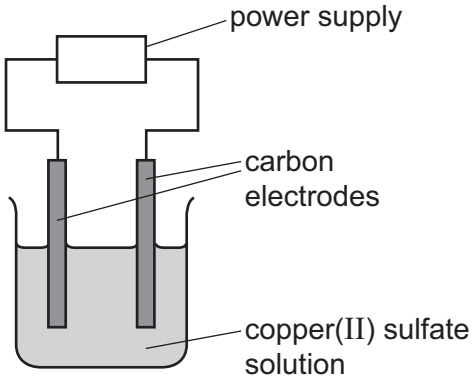
(iv) What difference would the student observe at the positive electrode if the aqueous copper(II) sulfate were replaced by concentrated aqueous copper(II) chloride?

..... [1]

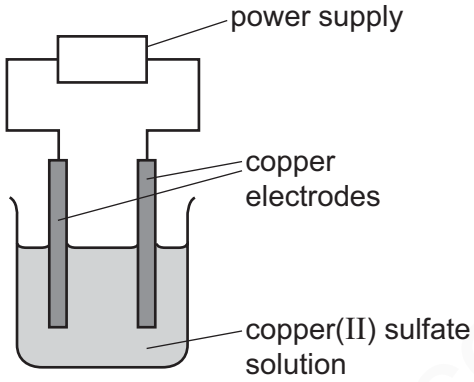
MARKING SCHEME:

(i)	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ 1 mark for any equation which has Cu as the product or Cu^{2+} ions on left 1 mark for correct species 1 mark for correct state symbols	**3**
(ii)	(a pink / brown) solid / deposit forms	**1**
(iii)	bubbles / fizzing (at the anode)	**1**
	solution becomes paler / less blue / colourless	**1**
(iv)	a green gas would be seen (on the anode)	**1**

2 A student electrolysed copper(II) sulfate solution using the two sets of apparatus shown.



apparatus **A**



apparatus **B**

In apparatus **A** the student used carbon electrodes.
In apparatus **B** the student used copper electrodes.

The student made the following observations.

apparatus **A**	apparatus **B**
The mass of the negative electrode increased.	The mass of the negative electrode increased.
The mass of the positive electrode stayed the same.	The mass of the positive electrode decreased.
Bubbles were seen at the positive electrode.	No bubbles were seen at the positive electrode.

(i) Explain why the mass of the negative electrode increased in **both** sets of apparatus.

.....

..... [1]

(ii) Name the gas that formed the bubbles seen in apparatus **A**.

..... [1]

(iii) Explain why the mass of the positive electrode decreased in apparatus **B**.

.....

..... [1]

(iv) Suggest what happens to the colour of the solution in apparatus **A** and apparatus **B** as the electrolysis progresses.
Explain your answer.

colour of the solution in apparatus **A** .....

colour of the solution in apparatus **B** .....

explanation .....

.....

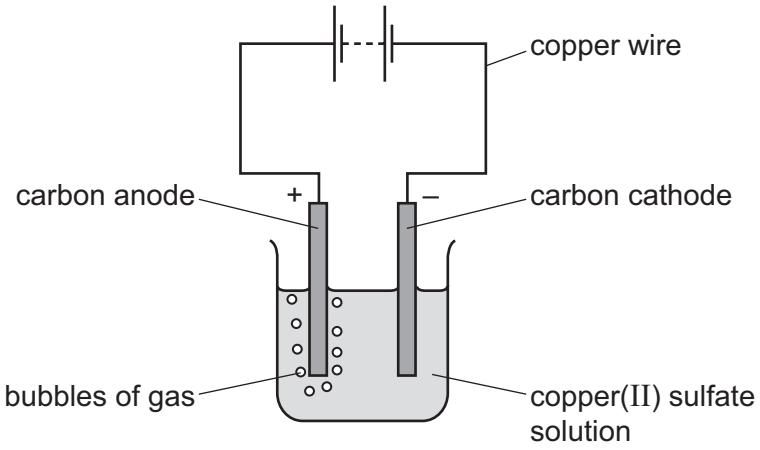
.....

[3]

MARKING SCHEME:

(i)	copper formed/copper deposited	**1**
(ii)	oxygen	**1**
(iii)	copper removed or copper lost **or** copper forms ions	**1**
(iv)	any **three** from: (apparatus **A**): solution becomes paler/fades in **A** (apparatus **B**): solution stays the same colour in **B** (explanation): copper ions removed (but not added) copper ions not replaced in **A** **OR** copper ions both removed and added (at the same rate) copper ions are being replaced (continually)	**3**

3 Copper(II) sulfate solution was electrolysed using the apparatus shown.



(a) A gas was formed at the anode.

Identify this gas and give the test for this gas.

gas .....

test .....

result of test .....

[3]

(b) During electrolysis, electricity passes through the copper(II) sulfate solution.

Solid copper(II) sulfate does not conduct electricity.

Explain **both** of these statements.

.....

.....

.....

.....

..... [3]

(c) The electrolysis was repeated using copper electrodes in place of carbon electrodes. The ionic half-equations for the reactions at the two electrodes are shown.

anode $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

cathode $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

(i) Which species is reduced during the electrolysis? Explain your answer.

.....

.....

..... [2]

(ii) The masses of the copper electrodes changed during the electrolysis.

State how **and** explain why the masses of the **two** copper electrodes changed.
Use the ionic half-equations to help you.

.....

.....

.....

.....

.....

.....

..... [3]

(iii) Explain why, during the electrolysis, the colour of the copper(II) sulfate solution does **not** change.

.....

.....

..... [1]

[Total: 12]

MARKING SCHEME:

(a)	(gas) oxygen (test) glowing splint (result of test) relights	**1** **1** **1**
(b)	reference to ions / ionic ions cannot move in solid **OR** are in fixed positions in solid ions can move when in solution	**1** **1** **1**
(c)(i)	copper ions / Cu^{2+} gain of electrons / oxidation number decreases	**1** **1**
(c)(ii)	any 3 from: anode decreases (in mass) copper removed (from anode) / solid (copper from anode) becomes aqueous cathode increases (in mass) copper deposited / added / Cu^{2+} deposited as Cu (on cathode)	**3**
(c)(iii)	copper is both added and removed (at same rate) **OR** the concentration (of copper ions) does not change	**1**

SMART EXAM RESOURCES
SUBJECT: CHEMISTRY
TOPIC:ELECTROCHEMISTRY
SET-3-QP-MS

ELECTROLYSIS OF COPPER SULFATE

This question is about electricity and chemical reactions.

1 **(a)** Aqueous copper(II) sulfate is an electrolyte.
The electrolysis of aqueous copper(II) sulfate using inert electrodes forms:

- copper at the cathode
- oxygen at the anode.

(i) State what is meant by the term electrolyte.

.....

..... [2]

(ii) State the term given to the Roman numeral, (II), in the name copper(II) sulfate.

..... [1]

(iii) State what happens to the colour of the aqueous copper(II) sulfate as this electrolysis progresses.

..... [1]

(iv) Write an ionic half-equation for the formation of copper at the cathode.

..... [2]

(v) Give the formula of the ion that forms oxygen at the anode.

..... [1]

(b) The electrolysis of aqueous copper(II) sulfate is repeated using **copper** electrodes.

State what happens to the anode.

..... [1]

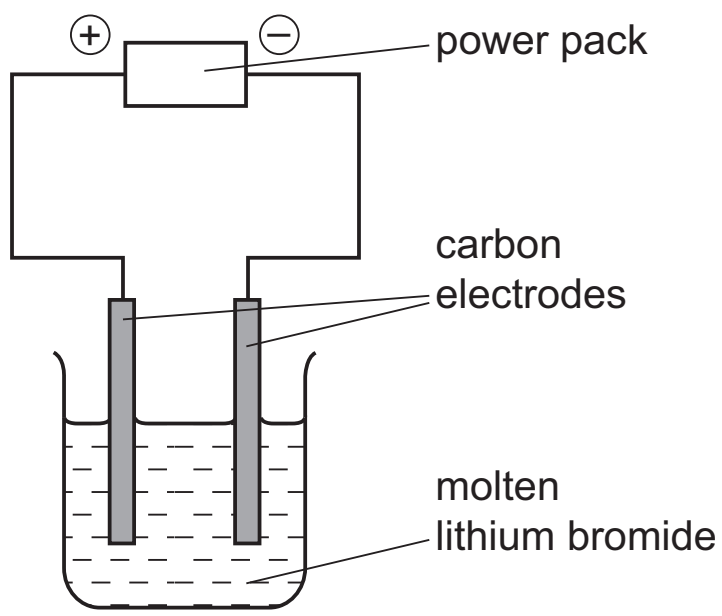
MARK SCHEME:

Question	Answer	Marks
(a)(i)	**M1** ionic compound **M2** molten and / or aqueous	**2**
(a)(ii)	oxidation number (of copper)	**1**
(a)(iii)	fades / (becomes) colourless	**1**
(a)(iv)	$Cu^{2+} + 2e \rightarrow Cu$ **M1** Cu^{2+} and (any number of) e on left hand side **M2** equation correct	**2**
(a)(v)	OH^-	**1**
(b)	anode dissolves	**1**

ELECTROLYSIS OF MOLTEN COMPOUNDS

5.6.1

The diagram shows the electrolysis of molten lithium bromide.



(i) Mark on the diagram the direction of the electron flow. [1]

(ii) Write an ionic equation for the reaction at the negative electrode (cathode).

..... [1]

(iii) Write an ionic equation for the reaction at the positive electrode (anode).

..... [2]

(iv) Which ion is oxidised? Explain your answer.

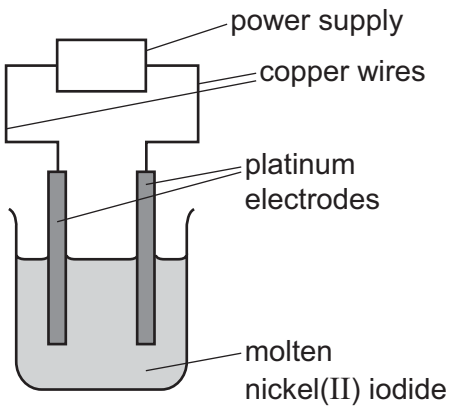
.....

..... [2]

(i)	correct direction (going towards negative electrode);	
(ii)	$Li^+ + e^- \rightarrow Li / Li^+ \rightarrow Li - e^-$;	
(ii	$2Br^- \rightarrow Br_2 + 2e^- / 2Br^- - 2e^- \rightarrow Br_2$ formulae; balancing;	
(iv)	Br^- / bromide (ion); electron lost / donated electrons / increased oxidation state / increased oxidation number / oxidation numbers changed from –1 to 0 / increased valency;	

ELECTROLYSIS OF MOLTEN COMPOUNDS

1 Molten nickel(II) iodide can be electrolysed using the apparatus shown.



During electrolysis, charge is transferred through the copper wires and through the molten nickel(II) iodide.

(i) Name the type of particles which transfer charge through the copper wires.

..... [1]

(ii) Name the type of particles which transfer charge through the molten nickel(II) iodide.

..... [1]

(iii) Predict the products of the electrolysis of molten nickel(II) iodide. Write an ionic half-equation for the formation of **one** of these products.

products .....

ionic half-equation .....

[3]

MARKING SCHEME:

(i)	electrons	**1**
(ii)	(positive and negative) ions	**1**
(iii)	nickel	**1**
	iodine	**1**
	$Ni^{2+} + 2e^- \rightarrow Ni$ **OR** $2\,I^- \rightarrow I_2 + 2\,e^-$	**1**

2 Molten sodium chloride is electrolysed using carbon electrodes.

(i) Name the product formed at the negative electrode.

..... [1]

(ii) Write an ionic half-equation for the reaction occurring at the negative electrode.

..... [1]

(iii) Chlorine is produced at the positive electrode.

Give the test for chlorine.

test .....

result .....

[2]

MARKING SCHEME:

(i)	sodium	1
(ii)	$Na^+ + e(^-) \rightarrow Na$	1
(iii)	*test:* (damp blue) litmus *result:* bleached / removes colour / (turns) white	1 1

SMART EXAM RESOURCES

SUBJECT: CHEMISTRY

TOPIC:ELECTROCHEMISTRY

SET-3-QP-MS

ELECTROLYSIS OF MOLTEN COMPOUNDS

1 Molten potassium chloride undergoes electrolysis.

(i) State what is meant by the term *electrolysis*.

.....

..... [2]

(ii) Name the products formed at the positive electrode (anode) and negative electrode (cathode) when molten potassium chloride undergoes electrolysis.

anode .....

cathode .....

[2]

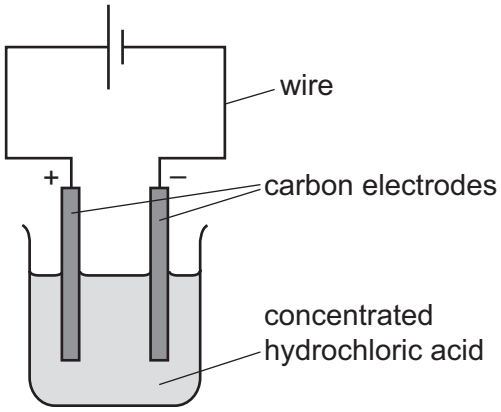
MARK SCHEME:

(i)	**breakdown** by (the passage of) **electricity** (1) of an **ionic compound** in **molten or aqueous** (state) (1)	**2**

(ii)	(anode) chlorine (cathode)potassium	**1**

2 This question is about electrolysis.

Concentrated hydrochloric acid is electrolysed using the apparatus shown.



The electrolysis is repeated using molten lead(II) bromide.

Describe what is seen at the:

- cathode .....
- anode. .....

[2]

MARK SCHEME:

M1 cathode: silver / grey solid (1) **M2** anode: bubbles of orange / brown gas (1)	**2**

3 The student carries out an electrolysis experiment on molten lead(II) chloride using the apparatus shown in Fig. 4.1. Chlorine gas forms at the anode and escapes from the apparatus.

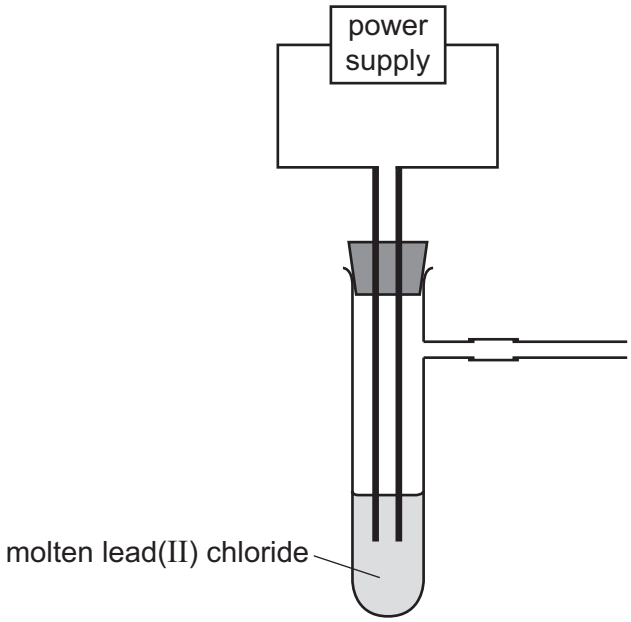


Fig. 4.1

(i) Explain why lead(II) chloride needs to be molten before it will conduct electricity.

.....

..... [1]

(ii) Write the ionic half-equation for the reaction occurring at the anode.

..... [2]

(iii) State the test for chlorine gas.

test .....

observations .....
[2]

(iv) Describe what is observed at the cathode.

..... [1]

(i)	mobile ions	**1**
(ii)	$2Cl^- \rightarrow Cl_2 + 2e^-$ **M1** any negative Cl species **losing** electron(s) (1) **M2** correct ionic half equation (1)	**2**
(iii)	**M1** (damp) litmus (paper) (1) **M2** is bleached / goes white (1)	**2**
iv	(shiny) grey **AND** solid	**1**

ELECTROLYSING BRINE

1 Hydrogen can also be manufactured by electrolysis. The electrolyte is concentrated aqueous sodium chloride. The electrodes are inert.

The products of electrolysis are hydrogen, chlorine and sodium hydroxide.

(i) Define the term *electrolysis*.

.....

..... [2]

(ii) Name a substance that can be used as the inert electrodes.

..... [1]

(iii) Write an ionic half-equation for the reaction in which hydrogen is produced.

..... [1]

(iv) Where is hydrogen produced in the electrolytic cell?

..... [1]

MARKING SCHEME:

(i)	**M1** breakdown of an ionic compound when molten or in aqueous solution; **M2** (using) electricity/electric current/electrical energy;	**2** 1 1
(ii)	carbon/graphite/platinum;	**1**
(iii)	$2H^+ + 2e(^-) \rightarrow H_2$; **OR** $2H_3O^+ + 2e(^-) \rightarrow H_2 + 2H_2O$;	**1**
(iv)	cathode/negative electrode;	**1**

2 **(a)** The electrolysis of concentrated aqueous sodium chloride can be represented by the following word equation.

sodium chloride + water → sodium hydroxide + hydrogen + chlorine

Construct a chemical equation to represent this reaction. Do not include state symbols.

..... [2]

(b) State one use of

chlorine, ...

sodium hydroxide, ...

hydrogen. .....

[3]

MARKING SCHEME:

(a)	$2NaCl + 2H_2O \rightarrow 2NaOH + H_2 + Cl_2$ all formulae correct; balancing;	**2**
(b)	**M1** *chlorine*: treating (drinking) water / treating water in swimming pools / kill bacteria in water / chlorination of water / (manufacture of) paper products / plastics / PVC / dyes / textiles / medicines / antiseptics / insecticides / herbicides / fungicides / solvents / paints / disinfectant / bleach / hydrochloric acid; **M2** *sodium hydroxide*: drain cleaner / oven cleaner / extraction of aluminium / purification of bauxite / (manufacture of) biodiesel / paper / soap / detergents / washing powder / textiles / dyes; **M3** *hydrogen*: fuel / rocket fuel / fuel cells / in welding / (manufacture of) ammonia / NH_3 / margarine / methanol / hydrochloric acid / refrigerants;	**3** 1 1 1

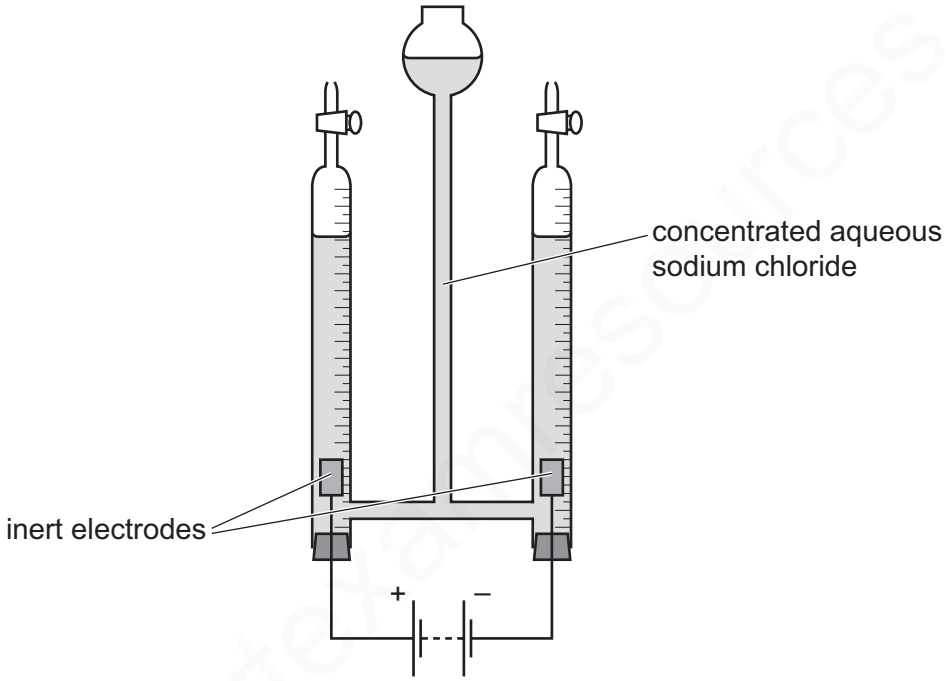
3 Many substances conduct electricity.

(a) Identify all the particles responsible for the passage of electricity in:

- graphite .....
- magnesium ribbon .....
- molten copper(II) bromide. .....

[4]

(b) A student used the following apparatus to electrolyse concentrated aqueous sodium chloride using inert electrodes.



(i) Suggest the name of a metal which could be used as the inert electrodes.

..... [1]

(ii) Name the gas formed at the positive electrode.

..... [1]

(iii) Write an ionic half-equation for the reaction occurring at the negative electrode. Include state symbols.

..... [3]

(iv) How, if at all, does the pH of the solution change during the electrolysis? Explain your answer.

.....

.....

..... [3]

MARKING SCHEME:

(a)	electrons (1) electrons (1) Cu^{2+} (ions) (1) Br^- (ions) (1)	**4**
(b)(i)	platinum	**1**
(b)(ii)	chlorine	**1**
(b)(iii)	$2H^+(aq) + 2e^- \rightarrow H_2(g)$ $H^+ + e^-$ on left hand side (1) rest of equation (1) state symbols of (aq) $\rightarrow$ (g) (1)	**3**
(b)(iv)	increases (sodium) hydroxide is formed (sodium) hydroxide is an alkali	**3**

4 A **concentrated** aqueous solution of sodium chloride is electrolysed using carbon electrodes.

(i) Name the products formed at the electrodes.

product at the positive electrode (anode) .....

product at the negative electrode (cathode) .....
[2]

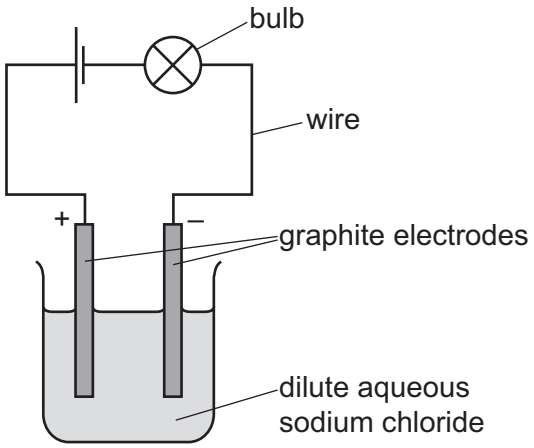
(ii) Write an ionic half-equation for the reaction occurring at the negative electrode.

..... [1]

MARKING SCHEME:

(i)	*product at the positive electrode:* chlorine *product at the negative electrode:* hydrogen	**1** **1**
(ii)	$2H^+ + 2e^- \rightarrow H_2$ **OR** $2H_3O^+ + 2e^- \rightarrow H_2 + 2H_2O$	**1**

5 A student sets up the following electrolysis experiment.



(a) Define the term *electrolysis*.

.....

..... [2]

(b) The student observes bubbles of colourless gas forming at each electrode.

(i) Name the main gas produced at the positive electrode (anode).

..... [1]

(ii) Describe a test for the gas produced in **(b)(i)**.

test .....

result .....
[2]

(iii) Write the ionic half-equation for the reaction taking place at the negative electrode (cathode).

..... [2]

(c) Charge is transferred during electrolysis.

Name the type of particle responsible for the transfer of charge in

the wires, ...

the electrolyte. .....
[2]

(d) The student replaces the dilute aqueous sodium chloride with **concentrated** aqueous sodium chloride.

Suggest **two** differences that the student observes.

1 .....

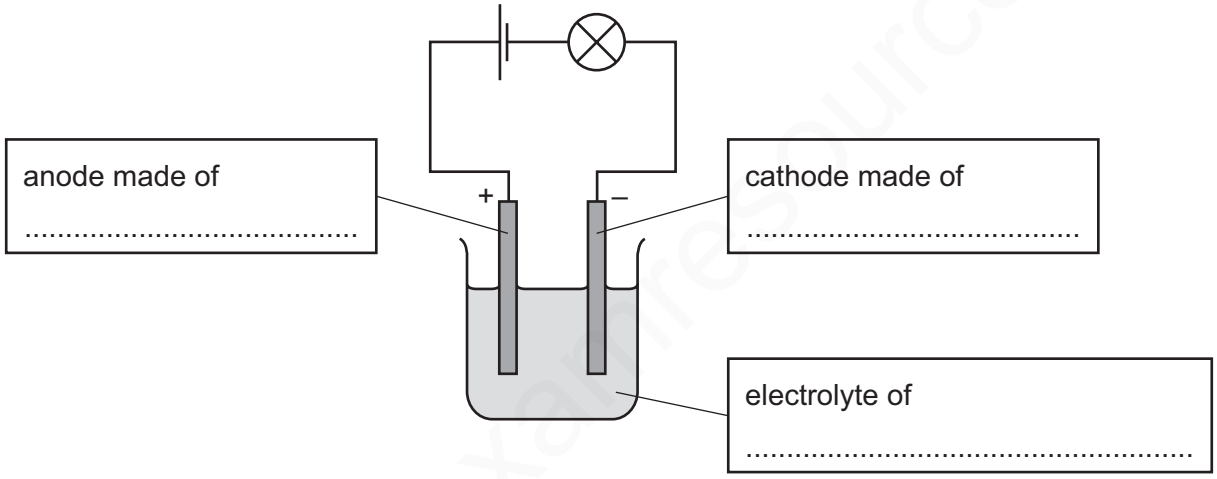
2 .....

[2]

(e) The student has a small piece of impure copper. The main impurities in the copper are small quantities of silver and zinc.

The student uses electrolysis to extract pure copper from the small piece of impure copper.

(i) Complete the labels on the diagram of the student's electrolysis experiment.



[3]

(ii) Use your knowledge of the reactivity series to suggest what happens to the silver and zinc impurities. Explain your answers.

silver impurities .....

.....

.....

zinc impurities .....

.....

.....

[3]

[Total: 17]

MARKING SCHEME:

(a)	the breakdown (into elements)	**1**
	of an (ionic) compound by (the passage of) electricity	**1**
(b)(i)	oxygen	**1**
(b)(ii)	**glowing** splint	**1**
	relights	**1**
(b)(iii)	$2H^+ + 2e^- \rightarrow H_2$ **M1** gain of electrons by H^+ **M2** rest of equation	**2**
(c)	*the wires:* electrons	**1**
	the electrolyte: ions	**1**
(d)	any 2 from: • green gas at positive electrode • bulb is brighter • rate of bubbles increases	**2**
(e)(i)	*anode made of:* impure copper	**1**
	cathode made of: (pure) copper	**1**
	electrolyte of: (aqueous) copper sulfate	**1**
(e)(ii)	silver (impurities) fall to the bottom of the cell	**1**
	zinc (impurities) (dissolve) into solution (as ions)	**1**
	because zinc is more reactive than copper **AND** silver is less reactive than copper	**1**

6 The electrolysis of concentrated aqueous sodium chloride, between inert electrodes, is used to make four important chemicals.

hydrogen
chlorine
sodium hydroxide
sodium chlorate(I)

(a) The ions present in the electrolyte are Na^+, H^+, Cl^- and OH^-.

(i) Hydrogen ions are discharged at the negative electrode (cathode).
Write an equation for this reaction.

..... [2]

(ii) The hydrogen ions are from the water.

$$H_2O \rightleftharpoons H^+ + OH^-$$

Suggest an explanation why the concentration of hydroxide ions increases.

.....

..... [2]

(iii) When a dilute solution of sodium chloride is used, chlorine is not formed at the positive electrode (anode), a different gas is produced. Name this gas.

..... [1]

(iv) State an example of an inert electrode.

..... [1]

(b) **(i)** State a use of hydrogen.

..... [1]

(ii) Why is chlorine used to treat the water supply?

..... [1]

MARKING SCHEME:

(a) (i) H_2 on RHS [1]
ignore any other species on RHS
rest of equation fully correct i.e. $2H^+ + 2e \rightarrow H_2$ [1]

(ii) $\underline{H^+}$ removed / escapes / discharged / used up / reduced [1]
(equilibrium) moves to RHS / more water molecules ionise or dissociate / forward reaction favoured [1]

(iii) oxygen / O_2 [1]
not O

(iv) carbon / graphite / platinum (electrode) [1]

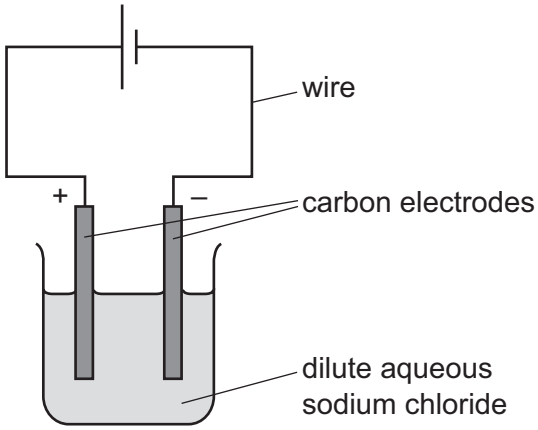
(b) (i) to make ammonia / in petroleum processing / balloons / rocket fuel / fuel for cars / hardening of fats / fuel cells / fuel (unqualified) / making hydrochloric acid [1]

(ii) to sterilise / disinfect it / kill bacteria / bugs / microbes / micro-organisms / germs [1]

SMART EXAM RESOURCES
SUBJECT: CHEMISTRY
TOPIC:ELECTROCHEMISTRY
SET-2-QP-MS

ELECTROLYSING BRINE

1 A student carries out an electrolysis experiment using the apparatus shown.



The student repeats the experiment using concentrated aqueous sodium chloride.

(i) Describe what the student observes at:

- the cathode .....
- the anode. .....

[2]

(ii) The student added litmus to the solution after the electrolysis of concentrated aqueous sodium chloride.

State the colour seen in the solution. Give a reason for your answer.

colour of solution .....

reason .....

[2]

(b) Carbon electrodes are used because they are inert.

State another element that can be used instead of carbon.

..... [1]

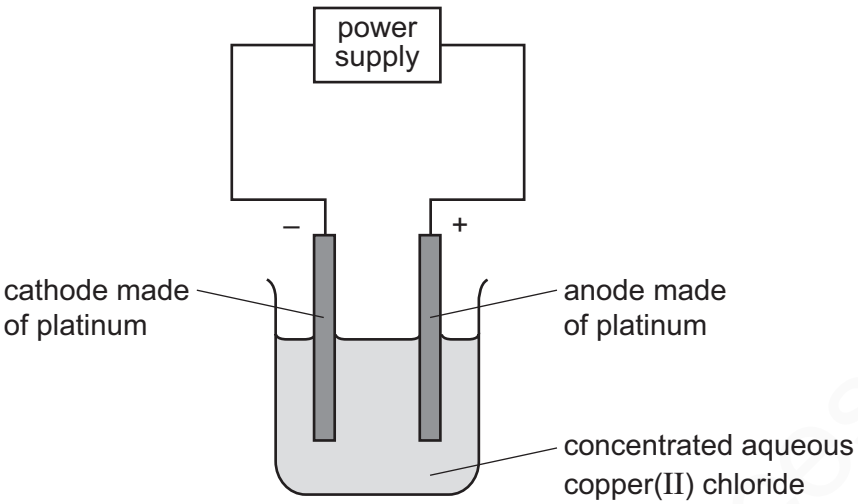
MARK SCHEME:

(a) (i)	fizzing (1) green gas (1)	**2**
(ii)	(litmus turns) blue **and** alkali / base forms (1) Sodium hydroxide / NaOH (forming) (1)	**2**
(b)	platinum	**1**

ELECTROLYSIS OF CONC SOLUTIONS

1 Solutions of ionic compounds can be broken down by electrolysis.

(a) Concentrated aqueous copper(II) chloride was electrolysed using the apparatus shown.



The ionic half-equations for the reactions at the electrodes are shown.

negative electrode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

positive electrode: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

(i) Platinum is a solid which is a good conductor of electricity.

State **one** other property of platinum which makes it suitable for use as electrodes.

.....

..... [1]

(ii) State what would be **seen** at the positive electrode during this electrolysis.

.....

..... [1]

(iii) State and explain what would happen to the mass of the negative electrode during this electrolysis.

.....

.....

..... [2]

(iv) The concentrated aqueous copper(II) chloride electrolyte is green.

Suggest what would happen to the colour of the electrolyte during this electrolysis.
Explain your answer.

.....

.....

..... [2]

(v) Identify the species that is oxidised during this electrolysis.
Explain your answer.

species that is oxidised .....

explanation .....

.....
[2]

MARKING SCHEME:

(a)(i)	inert / unreactive / does not react with chlorine	**1**
(a)(ii)	bubbles / fizzing / effervescence	**1**
(a)(iii)	**M1** increases **M2** (solid) copper deposited	**2**
(a)(iv)	**M1** colour fades / becomes pale(r) / becomes colourless / becomes lighter **M2** copper (ions) removed (from solution)	**2**
(a)(v)	**M1** species oxidised: chloride (ions) / Cl^- **M2** explanation: loss of electrons / increase in oxidation state	**2**

2 This question is about electrolysis.

(a) (i) What is meant by the term *electrolysis*?

.....

..... [2]

(ii) Name the type of particle responsible for the conduction of electricity during electrolysis in:

the metal wires .....

the electrolyte .....

[2]

(b) The table gives information about the products of the electrolysis of two electrolytes. Platinum electrodes are used in each case.

(i) Give **two** reasons why platinum is suitable to use as an electrode.

1 .....

2 .....

[2]

(ii) Complete the table.

electrolyte	observation at the anode (+)	name of product at the anode (+)	observation at the cathode (–)	name of product at the cathode (–)
concentrated aqueous potassium chloride			bubbles of colourless gas	
aqueous copper(II) sulfate	bubbles of colourless gas			

[6]

[Total: 12]

MARKING SCHEME:

<table>
<tr><td>(a)(i)</td><td>M1 breakdown of an ionic compound when molten or in aqueous solution
M2 (using) electricity / electric current</td><td>2</td></tr>
<tr><td>(a)(ii)</td><td>M1 electron(s)
M2 ion(s)</td><td>2</td></tr>
<tr><td>(b)(i)</td><td>M1 inert / unreactive
M2 conducts electricity</td><td>2</td></tr>
<tr><td>(b)(ii)</td><td>
<table>
<tr><td>observation at anode(+)</td><td>name of product at anode(+)</td><td>observation at cathode(–)</td><td>name of product at cathode(–)</td></tr>
<tr><td>M1 green / yellow bubbles</td><td>M2 chlorine</td><td></td><td>M3 hydrogen</td></tr>
<tr><td></td><td>M4 oxygen</td><td>M5 pink / brown solid</td><td>M6 copper</td></tr>
</table>
</td><td>6</td></tr>
</table>

3 A sample of concentrated hydrobromic acid, HBr(aq), was electrolysed using platinum electrodes.
The concentration of the hydrobromic acid was 8.89 mol / dm^3.

(i) Calculate the concentration of the HBr(aq) in g / dm^3.

concentration of HBr(aq) = ... g / dm^3 [1]

(ii) Explain why concentrated HBr(aq) can conduct electricity.

.....

.....

..... [2]

(iii) Magnesium is **not** a suitable material from which to make the electrodes.

Explain why.

.....

..... [1]

(iv) Predict the product formed at the anode when concentrated HBr(aq) is electrolysed.

..... [1]

(v) Write the ionic half-equation for the reaction occurring at the cathode.

..... [2]

MARKING SCHEME:

(i)	720(.09)	**1**
(ii)	(it contains) ions (1) (ions) are able to move (1)	**2**
(iii)	magnesium is not inert	**1**
(iv)	bromine / Br_2	**1**
(v)	H^+ and $e^{(-)}$ on LHS (1) fully correct, i.e.: $2H^+ + 2e^- \rightarrow H_2$ (1)	**2**

ELECTROLYSIS -DILUTE NaCl

1 (a) A **dilute** aqueous solution of sodium chloride is electrolysed using carbon electrodes.

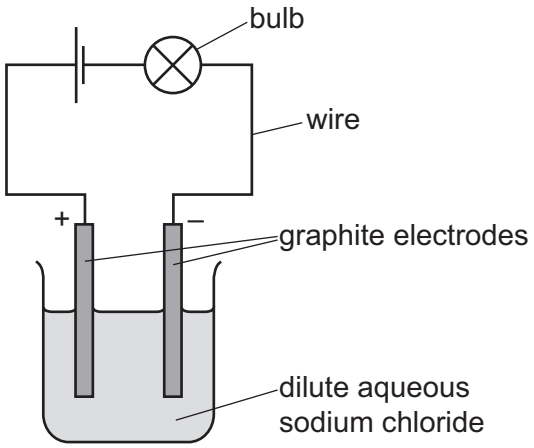
Name the main product formed at the positive electrode.

..... [1]

MARKING SCHEME:

(a)	oxygen	1

2 A student sets up the following electrolysis experiment.



(a) Define the term *electrolysis*.

.....

..... [2]

(b) The student observes bubbles of colourless gas forming at each electrode.

(i) Name the main gas produced at the positive electrode (anode).

..... [1]

(ii) Describe a test for the gas produced in **(b)(i)**.

test .....

result .....

[2]

(iii) Write the ionic half-equation for the reaction taking place at the negative electrode (cathode).

..... [2]

(c) Charge is transferred during electrolysis.

Name the type of particle responsible for the transfer of charge in

the wires, ...

the electrolyte. .....

[2]

(d) The student replaces the dilute aqueous sodium chloride with **concentrated** aqueous sodium chloride.

Suggest **two** differences that the student observes.

1 .....

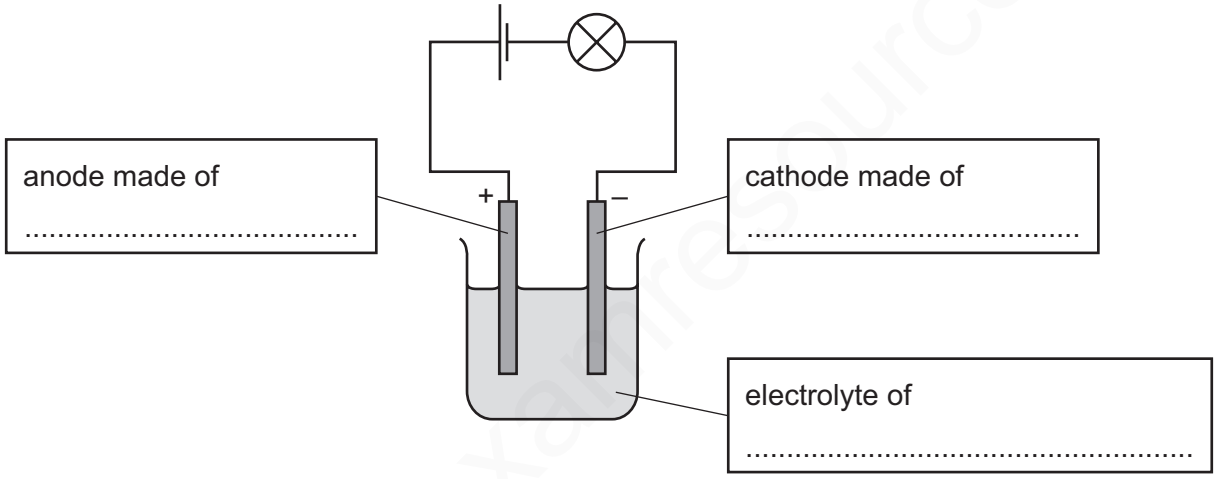
2 .....

[2]

(e) The student has a small piece of impure copper. The main impurities in the copper are small quantities of silver and zinc.

The student uses electrolysis to extract pure copper from the small piece of impure copper.

(i) Complete the labels on the diagram of the student's electrolysis experiment.



[3]

(ii) Use your knowledge of the reactivity series to suggest what happens to the silver and zinc impurities. Explain your answers.

silver impurities .....

.....

.....

zinc impurities .....

.....

.....

[3]

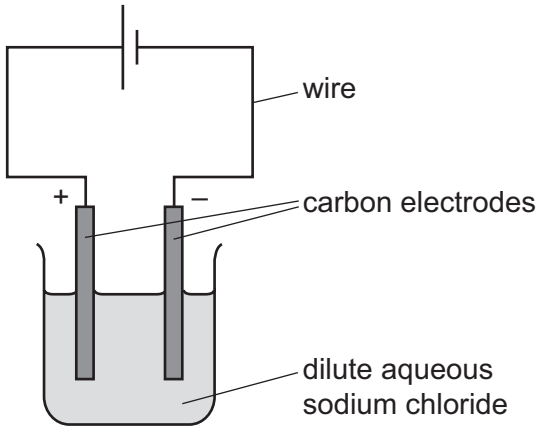
[Total: 17]

MARKING SCHEME:

(a)	the breakdown (into elements)	**1**
	of an (ionic) compound by (the passage of) electricity	**1**
(b)(i)	oxygen	**1**
(b)(ii)	**glowing** splint	**1**
	relights	**1**
(b)(iii)	$2H^+ + 2e^- \rightarrow H_2$ **M1** gain of electrons by H^+ **M2** rest of equation	**2**
(c)	*the wires:* electrons	**1**
	the electrolyte: ions	**1**
(d)	any 2 from: • green gas at positive electrode • bulb is brighter • rate of bubbles increases	**2**
(e)(i)	*anode made of:* impure copper	**1**
	cathode made of: (pure) copper	**1**
	electrolyte of: (aqueous) copper sulfate	**1**
(e)(ii)	silver (impurities) fall to the bottom of the cell	**1**
	zinc (impurities) (dissolve) into solution (as ions)	**1**
	because zinc is more reactive than copper **AND** silver is less reactive than copper	**1**

ELECTROLYSIS OF DILUET NaCl

1 A student carries out an electrolysis experiment using the apparatus shown.



The student uses dilute aqueous sodium chloride.

(a) State the name given to any solution which undergoes electrolysis.

..... [1]

(b) Hydroxide ions are discharged at the anode.

(i) Complete the ionic half-equation for this reaction.

.....$OH^-(aq)$ → ..... + $O_2(g)$ + $4e^-$ [2]

(ii) Explain how the ionic half-equation shows the hydroxide ions are being oxidised.

..... [1]

(c) Describe what the student observes at the cathode.

..... [1]

(d) Write the ionic half-equation for the reaction at the cathode.

..... [2]

MARK SCHEME:

(a)	electrolyte	1
(b)(i)	$4OH^- \rightarrow 2H_2O + O_2 + 4e^-$ balance of charge (1) rest of equation (1)	2
(b)(ii)	(OH^-(aq) ions) lose electrons	1
(c)	fizzing	1
(d)	$2H^+ + 2e^- \rightarrow H_2$ species correct (1) fully correct equation (1)	2

USE OF CHLORINE

5.4.1

Why does the water supply industry use chlorine?

..... [1]

----------**Marking Scheme**----------

sterilise/disinfect water **or** kill microbes/germs bacteria, etc. [1]
NOT just to make it safe to drink **or** purify it **or** clean it
treat above as neutral they do not negate a correct response

5.4.2

Describe an industrial method of making chlorine.

.....

..... [2]

electrolysis [1]
aqueous sodium chloride [1]

INERT ELECTRODES

1 The results of experiments on electrolysis using inert electrodes are given in the table.

Complete the table; the first line has been completed as an example.

electrolyte	change at negative electrode	change at positive electrode	change to electrolyte
molten lead(II) bromide	lead formed	bromine formed	used up
....	potassium formed	iodine formed	used up
dilute aqueous sodium chloride	.....	.....	....
aqueous copper(II) sulfate	.....	.....	....
....	hydrogen formed	bromine formed	potassium hydroxide formed

[Total: 8]

MARKING SCHEME:

molten potassium iodide	**NOT** aqueous	[1]
hydrogen		[1]
oxygen		[1]
water used up **or** solution becomes more concentrated **or** sodium chloride remains **NOT** no change		[1]
If products are given as hydrogen, chlorine and sodium hydroxide then 2/3		
copper		[1]
oxygen (and water)		[1]
sulfuric acid	accept hydrogen sulfate	[1]
aqueous **or** dilute **or** concentrated potassium bromide **accept** correct formulae		[1]

[Total: 8]

PURIFYING METALS

5.11.1

Suggest another method of refining nickel. Give a brief description of the method which you have suggested. A labelled diagram is acceptable.

[4]

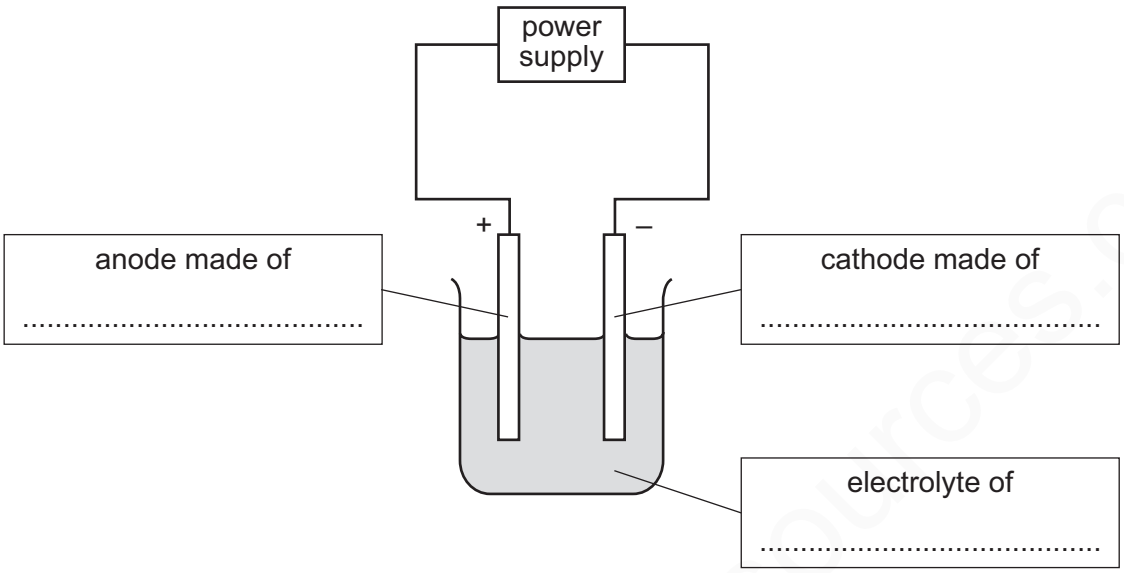
electrolysis	[1]
cathode (pure) nickel	[1]
anode impure nickel	[1]
electrolyte is a soluble nickel salt	[1]

PURIFICATION OF COPPER

1 Copper is refined (purified) by electrolysis. Nickel can be refined using a similar method.

(i) The diagram shows the refining of nickel by electrolysis.

Complete the labels in the boxes.



[3]

(ii) Indicate, by writing **N** on the diagram, where nickel is produced. [1]

MARKING SCHEME:

(i)	**M1** electrolyte aqueous or solution of named nickel salt (1) **M2 anode** impure nickel (1) **M3 cathode** pure nickel (1)	**3**
(ii)	nickel produced at cathode under the liquid surface (1)	**1**

2 Copper is refined by the electrolysis of aqueous copper(II) sulphate using copper electrodes. Describe the change that occurs at the electrodes.

(i) cathode (pure copper) .....

.....[1]

(ii) anode (impure copper) .....

.....[1]

(iii) Write an ionic equation for the reaction at the cathode.

.....[1]

(iv) If carbon electrodes are used, a colourless gas is given off at the anode and the electrolyte changes from a blue to a colourless solution.

The colourless gas is ..... .

The solution changes into ..... . [2]

MARKING SCHEME:

(i) copper deposited **or** mass increases [1]

(ii) copper goes into solution **or** mass decreases [1]

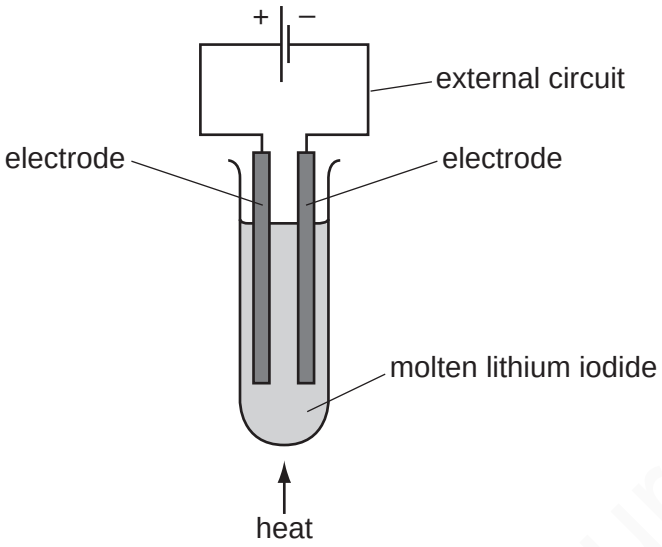
(iii) $Cu^{2+} + 2e^- \Longrightarrow Cu$ [1]

(iv) oxygen [1]
sulphuric acid accept hydrogen sulphate [1]

ELECTROLYSIS-MIXED BAG

1 During electrolysis, ions move in the electrolyte and electrons move in the external circuit. Reactions occur at the electrodes.

(a) The diagram shows the electrolysis of molten lithium iodide.



(i) Draw an arrow on the diagram to show the direction of the electron flow in the external circuit. [1]

(ii) Electrons are supplied to the external circuit. How and where is this done?

.....

..... [2]

(iii) Explain why solid lithium iodide does not conduct electricity but when molten it is a good conductor.

.....

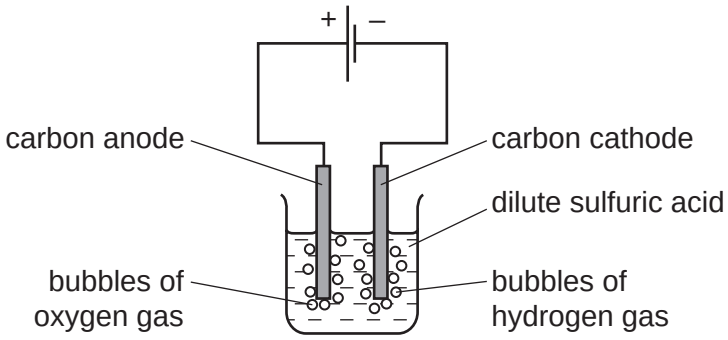
..... [1]

(b) The results of experiments on electrolysis are shown in the following table. Complete the table. The first line has been done as an example.

electrolyte	electrodes	product at cathode	product at anode	change to electrolyte
molten lithium iodide	carbon	lithium	iodine	used up
aqueous copper(II) sulfate	platinum		oxygen	
concentrated aqueous potassium chloride	carbon		chlorine	

[4]

(c) The diagram below shows the electrolysis of dilute sulfuric acid. Hydrogen is formed at the negative electrode (cathode) and oxygen at the positive electrode (anode) and the concentration of sulfuric acid increases.



The ions present in the dilute acid are $H^+(aq)$, $OH^-(aq)$ and $SO_4^{2-}(aq)$.

(i) Write an equation for the reaction at the negative electrode (cathode).

..... [2]

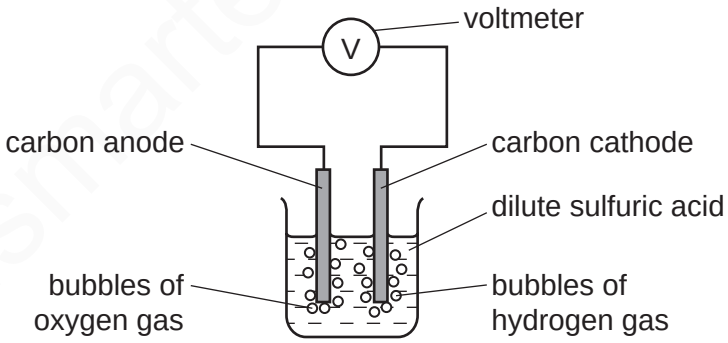
(ii) Complete the equation for the reaction at the positive electrode (anode).

$$4OH^-(aq) \rightarrow O_2(g) + ...H_2O(l) + ...$$ [1]

(iii) Suggest an explanation of why the concentration of the sulfuric acid increases.

..... [1]

(d) In the apparatus used in **(c)**, the power supply is removed and immediately replaced by a voltmeter.



A reading on the voltmeter shows that electrical energy is being produced. Suggest an explanation for how this energy is produced.

.....

.....

..... [3]

[Total: 15]

MARKING SCHEME:

(a) (i) correct arrow from negative terminal of battery or from anode; [1]

(ii) from battery / power supply / cell; [1]
from negative electrode of battery to external circuit; [1]
or from anode;
from iodide ion losing electron **or** oxidation of anion;

(iii) ions cannot move in solid / ions can move in liquid; [1]

(b) copper; [1]
(changes to) sulfuric acid; [1]

hydrogen; [1]
(changes to) potassium hydroxide; [1]

(c) (i) $2H^+ + 2e \rightarrow H_2$ [2]
not balanced = [1]

(ii) $4OH^- \rightarrow O_2 + 2H_2O + 4e$ [1]

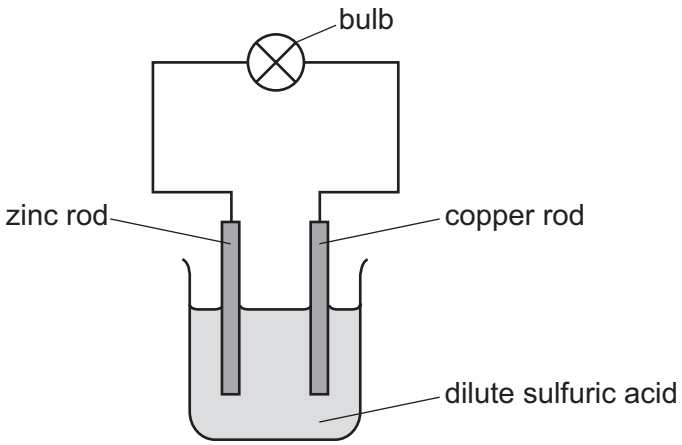
(iii) water used up; [1]

(d) it is a cell; [1]
hydrogen reacts with oxygen; [1]
this reaction produces energy / is exothermic / produces flow of electrons /
changes chemical energy to electrical energy; [1]

[Total: 15]

SIMPLE CELLS

1 **(a)** When rods of zinc and copper are placed into dilute sulfuric acid as shown, electricity is generated.



(i) Write the ionic half-equation for the reaction occurring at the zinc rod.

..... [2]

(ii) Write the ionic half-equation for the reaction occurring at the copper rod.

..... [2]

(iii) The copper rod was replaced by an iron rod.

Suggest the change, if any, in the intensity of the light emitted from the bulb and give a reason for your answer.

change .....

reason .....

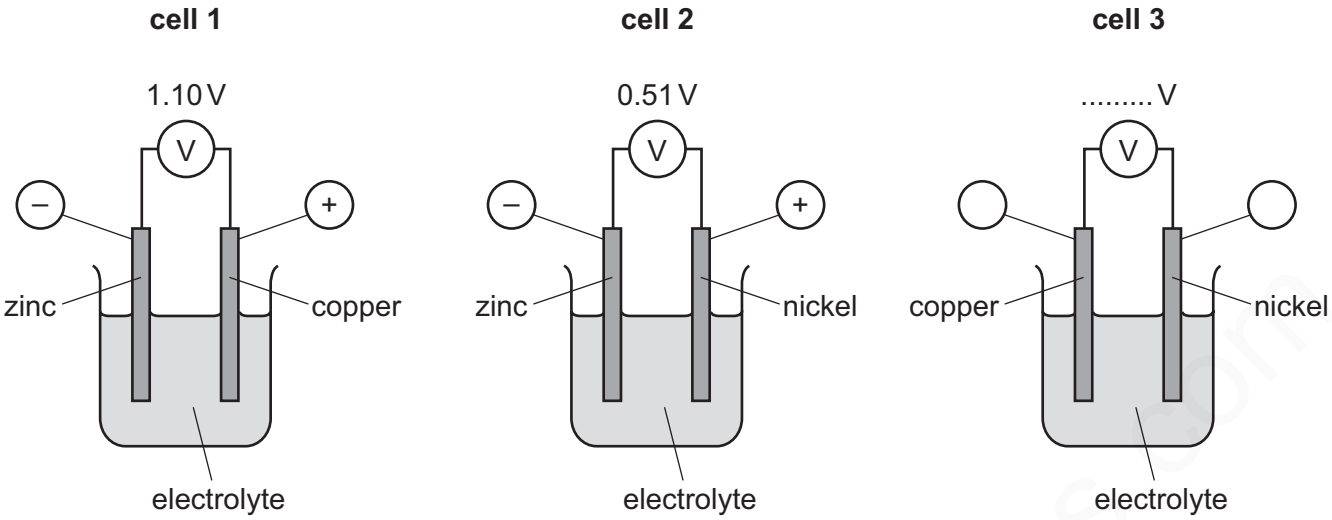
.....

[2]

MARKING SCHEME:

(i)	$Zn \rightarrow Zn^{2+} + 2e^-$ **M1** correct species **M2** correct balancing	**2**
(ii)	$2H^+ + 2e^- \rightarrow H_2$ **M1** correct species **M2** correct balancing	**2**
(iii)	change: (the intensity would) decrease	**1**
	reason: the difference in reactivity between zinc and iron is less than the difference in reactivity between zinc and copper	**1**

2 **(a)** Three cells are set up each using two metals.



(i) Write the ionic half-equation for the reaction occurring at the zinc electrode in **cell 1**.

..... [2]

(ii) Put the **three** metals, copper, nickel and zinc, in order of reactivity.

most reactive .....

↓ .....

least reactive .....

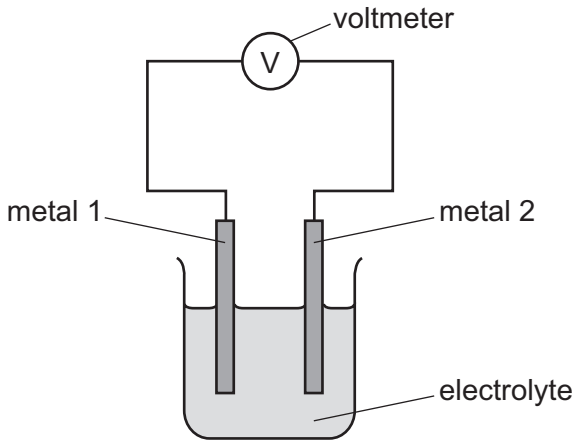
[1]

(iii) Complete the labelling in **cell 3** by writing the polarity (+/–) of each electrode in the circles and calculating the reading on the voltmeter. [2]

MARKING SCHEME:

(i)	$Zn \rightarrow Zn^{2+} + 2e / 2e^-$ **M1** formula of Zn^{2+} on the right-hand side **M2** equation fully correct	**2**
(ii)	zinc / Zn nickel / Ni copper / Cu	**1**
(iii)	copper (+) and nickel (–)	**1**
	0.59 V	**1**

3 The diagram shows a simple cell.



The simple cell was used with different metals as electrodes. The voltages were recorded in the table.

- If the voltage measured is positive then metal 2 is more reactive than metal 1.
- If the voltage measured is negative then metal 1 is more reactive than metal 2.

		metal 2				
		beryllium	cobalt	nickel	silver	vanadium
metal 1	beryllium	0.0 V	–1.6 V	–1.6 V	not measured	–0.7 V
	cobalt		0.0 V	0.0 V	–1.1 V	0.9 V
	nickel			0.0 V	–1.1 V	0.9 V
	silver				0.0 V	2.0 V
	vanadium					0.0 V

- The more reactive metal is oxidised.
- The bigger the difference in reactivity of the metals, the larger the reading on the voltmeter.

(a) In a simple cell using nickel and silver, the nickel is oxidised.

(i) Define *oxidation* in terms of electrons.

..... [1]

(ii) Nickel forms ions with a charge of +2.

Write an ionic half-equation to show the oxidation of nickel.

..... [1]

(iii) What will happen to the mass of the nickel electrode when the nickel is oxidised?

.....

..... [1]

(b) Use the data in the table to answer the following questions.

(i) Which of the metals in the table is the most reactive?
Explain your answer.

.....

..... [2]

(ii) State which **two** different metals have the same reactivity.

..... [1]

(iii) Predict the voltage produced by a simple cell with beryllium as metal 1 and silver as metal 2.

.....

..... [2]

(c) Describe how the simple cell in the diagram can be used to show that magnesium is more reactive than beryllium. Explain your answer.

.....

.....

.....

..... [2]

[Total: 10]

MARKING SCHEME:

(a)(i)	loss (of electrons)	**1**
(a)(ii)	$Ni \rightarrow Ni^{2+} + 2e^-$	**1**
(a)(iii)	goes down / gets less / decreases / lower / smaller	**1**
(b)(i)	beryllium	**1**
	most negative voltage with any (named) metal **OR** biggest voltage with cobalt / nickel	**1**
(b)(ii)	cobalt **AND** nickel	**1**
(b)(iii)	– sign	**1**
	2.7	**1**
(c)	(set up cell) using magnesium and beryllium (electrodes)	**1**
	voltage positive if magnesium is metal 2	**1**
	OR	
	(set up cells) using both magnesium and beryllium with the same metal as the other electrode	**1**
	larger (magnitude) voltages with magnesium	**1**
	OR	
	use magnesium with a different metal and compare to a reference value in a table	**1**
	value is more negative than with beryllium, if magnesium is metal 1	**1**

4 In the following list of ionic equations, the metals are in order of reactivity.

$\square \rightarrow \square + \square$

$Zn \rightarrow Zn^{2+} + 2e^-$

$Sn \rightarrow Sn^{2+} + 2e^-$

$Hg \rightarrow Hg^{2+} + 2e^-$

$Ag \rightarrow Ag^{+} + e^-$

↑ reactivity of metals increases

(a) (i) In the space at the top of the series, write an ionic equation that includes a more reactive metal. [1]

(ii) Define *oxidation* in terms of electron transfer.

.....

[1]

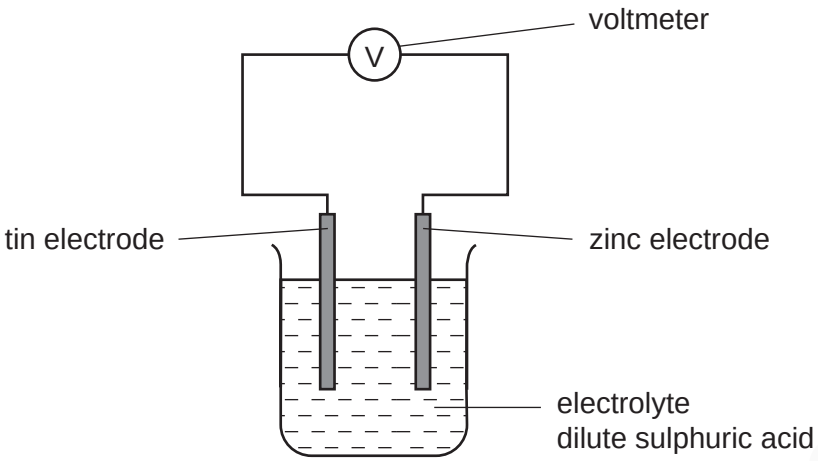
(iii) Explain why the positive ions are likely to be oxidising agents.

[1]

(iv) Which positive ion(s) can oxidise mercury metal (Hg)?

[1]

(b) The following diagram shows a simple cell.



(i) Predict how the voltage of the cell would change if the tin electrode was replaced with a silver one.

[1]

(ii) Which electrode would go into the solution as positive ions? Give a reason for your choice.

[1]

(iii) State how you can predict the direction of the electron flow in cells of this type.

[1]

MARKING SCHEME:

(a) (i) Correct equation with a more reactive metal **[1]**

(ii) Electron loss **[1]**

(iii) Because they can accept electrons or take electrons away from……. **[1]**

(iv) Silver or silver(I) **[1]**

(b) (i) increase **[1]**

(ii) zinc
COND and a correct reason - such as it loses electrons more easily **or** it is more reactive **[1]**
Need both zinc and reason for the mark.

(iii)from the more reactive to the less reactive **NOT** just from zinc to lead **[1]**

TOTAL = 7

5 One way of establishing a reactivity series is by displacement reactions.

(a) A series of experiments was carried out using the metals lead, magnesium, zinc and silver. Each metal was added in turn to aqueous solutions of the metal nitrates.

The order of reactivity was found to be:

magnesium — most reactive
zinc
↓
lead
silver — least reactive

(i) Complete the table.

✓ = reacts
✗ = does not react

	metal			
aqueous solution	lead Pb	magnesium Mg	zinc Zn	silver Ag
lead(II) nitrate		✓	✓	✗
magnesium nitrate				
zinc nitrate				
silver nitrate				

[3]

(ii) Displacement reactions are redox reactions.
On the following equation, draw a **ring** around the reducing agent and an **arrow** to show the change which is oxidation.

$$Zn + Pb^{2+} \rightarrow Zn^{2+} + Pb$$

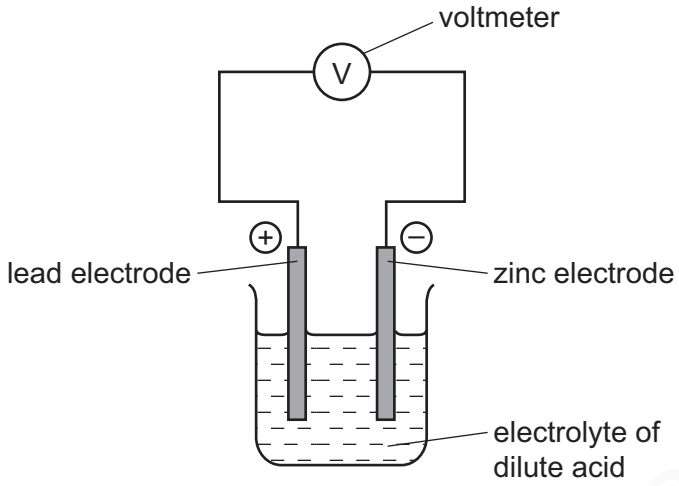
[2]

(iii) Complete the following ionic equation.

$$Zn + 2Ag^{+} \rightarrow \text{.....} + \text{.....}$$

[1]

(b) Another way of determining the order of reactivity of metals is by measuring the voltage and polarity of simple cells. The polarity of a cell is shown by which metal is the positive electrode and which metal is the negative electrode. An example of a simple cell is shown below.



(i) Mark on the above diagram the direction of the electron flow. [1]

(ii) Explain, in terms of electron transfer, why the more reactive metal is always the negative electrode.

.....

.....

..... [2]

(iii) The following table gives the polarity of cells using the metals zinc, lead, copper and manganese.

cell	electrode 1	polarity	electrode 2	polarity
A	zinc	–	lead	+
B	manganese	–	lead	+
C	copper	+	lead	–

What information about the order of reactivity of these four metals can be deduced from the table?

.....

.....

..... [2]

(iv) What additional information is needed to establish the order of reactivity of these four metals using cells?

..... [1]

[Total: 12]

MARKING SCHEME:

(a) (i)

aqueous solution	lead Pb	magnesium Mg	zinc Zn	silver Ag
lead (II) nitrate		✓	✓	✗
magnesium nitrate	✗		✗	✗
zinc nitrate	✗	✓		✗
silver(I) nitrate	✓	✓	✓	

each horizontal line correct (1) [3]

(ii) Zn (1)

An arrow **from** Zn **to** Zn^{2+} (1) [2]

(iii) $Zn + 2Ag^{+} \rightarrow Zn^{2+} + 2Ag$ (1) [1]

(b) (i) correct direction from zinc to lead (1) [1]

(ii) metals react by **losing electrons** (1)

the more reactive metal/zinc will lose electrons more readily (making the electrode negatively charged). (1) [2]

(iii) manganese **and** zinc are more reactive than lead (and/or copper) (1)

lead is more reactive than copper (1) [2]

(iv) the **polarity** of a Mn/Zn (cell)
or the **voltages** of Zn/Pb **and** Mn/Pb (cells) (1) [1]

[Total: 12]

SIMPLE CELLS

1 The reactivity series shows the metals in order of reactivity.

(a) The reactivity series can be established using displacement reactions. A piece of zinc is added to aqueous lead nitrate. The zinc becomes coated with a black deposit of lead.

$$Zn + Pb^{2+} \rightarrow Zn^{2+} + Pb$$

Zinc is more reactive than lead.

The reactivity series can be written as a list of ionic equations.

.....	→	.....	+	.....	most reactive metal : the best reductant (reducing agent)
Zn	→	Zn^{2+}	+	$2e^-$	
Fe	→	Fe^{2+}	+	$2e^-$	
Pb	→	Pb^{2+}	+	$2e^-$	
Cu	→	Cu^{2+}	+	$2e^-$	
Ag	→	Ag^+	+	e^-	

(i) In the space at the top of the list, write an ionic equation for a metal which is more reactive than zinc. [1]

(ii) Write an ionic equation for the reaction between aqueous silver(I) nitrate and zinc.

..... [2]

(iii) Explain why the positive ions are likely to be oxidants (oxidising agents).

..... [1]

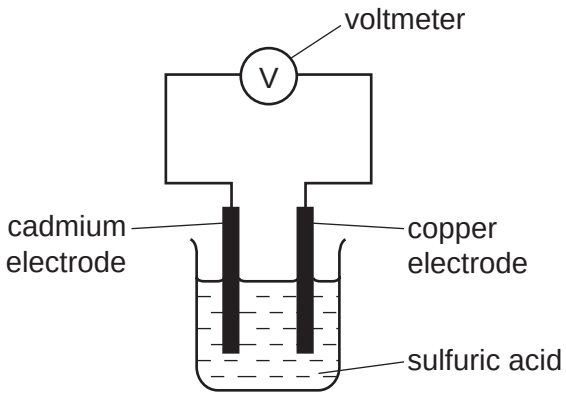
(iv) Deduce which ion is the best oxidant (oxidising agent).

..... [1]

(v) Which ion(s) in the list can oxidise lead metal?

..... [1]

(b) A reactivity series can also be established by measuring the voltage of simple cells. The diagram shows a simple cell.



Results from cells using the metals tin, cadmium, zinc and copper are given in the table below.

cell	electrode 1 positive electrode	electrode 2 negative electrode	voltage / volts
1	copper	cadmium	0.74
2	copper	tin	0.48
3	copper	zinc	1.10

Write the four metals in order of increasing reactivity and explain how you used the data in the table to determine this order.

.....

.....

..... [3]

[Total: 9]

MARKING SCHEME:

(a) **(i)** any metal above zinc
$Mg \rightarrow Mg^{2+} + 2e^-$ [1]

(ii) $Zn + 2Ag^+ \rightarrow Zn^{2+} + 2Ag$ [2]
Note: not balanced only [1]

(iii) because they can accept or gain electrons / change into atoms or can be reduced [1]

(iv) Ag^+ or silver [1]
charge not essential but if given must be correct

(v) Ag^+ and Cu^{2+} **or** silver and copper [1]
charge not essential but if given must be correct

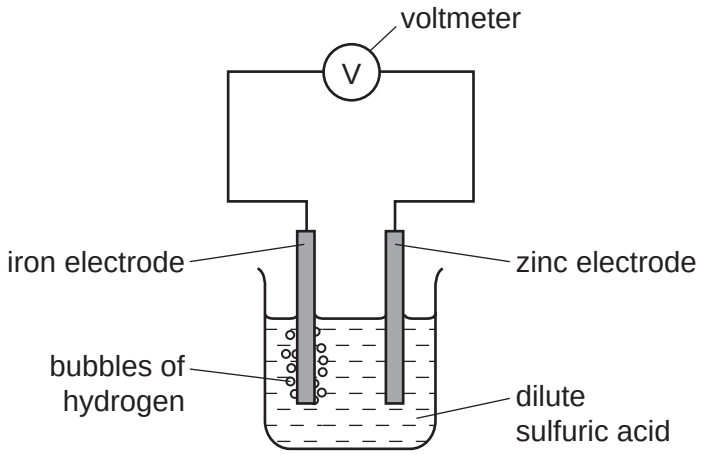
(b) Cu Sn Cd Zn *(i.e. all **4** in correct order)* [1]
relates order to voltage [1]

one relevant comment from: [1]

higher reactivity metals are the negative electrode / copper is least reactive because it is the positive electrode because copper would have the lowest voltage / copper cell V = 0 / the bigger the difference in reactivity, the bigger the voltage / zinc has highest voltage because it is most reactive / more reactive metals have higher voltage

[Total: 9]

2 The diagram shows a simple cell.



(a) Write an equation for the overall reaction occurring in the cell.

..... [2]

(b) Explain why all cell reactions are exothermic and redox.

.....

.....

..... [3]

(c) Which electrode, zinc or iron, is the negative electrode? Give a reason for your choice.

.....

..... [2]

(d) Suggest **two** ways of increasing the voltage of this cell.

.....

..... [2]

[Total: 9]

MARKING SCHEME:

(a) $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$ / $Zn + 2H^+ \rightarrow Zn^{2+} + H_2$ [2]

marks are for correct reactants [1] correct products [1]
If ionic equation is given don't penalise SO_4^{2-} spectator ions on both sides

(b) (exothermic because) a cell produces (electrical) energy/electricity [1]

the next two marks score for

electrons are lost **AND** gained / oxidation no. or state/valency **both** increases and decreases / two correct half equations i.e. $Zn \rightarrow Zn^{2+} + 2e^-$ and $2H^+ + 2e^- \rightarrow H_2$ [2]

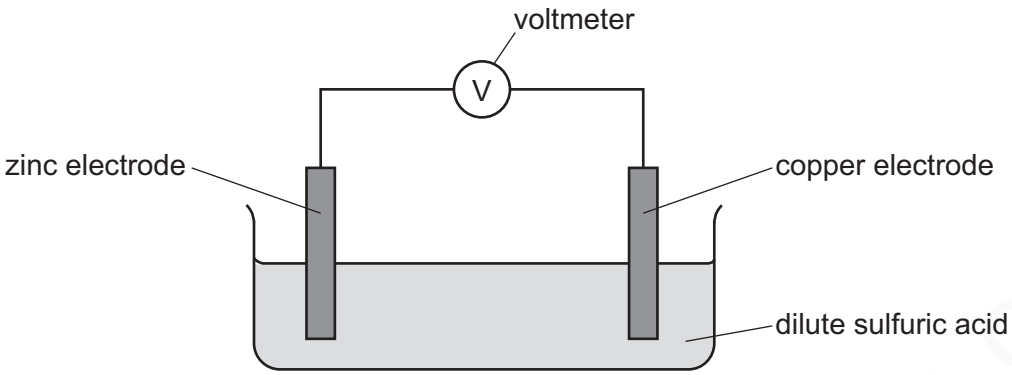
(c) zinc [1]
cond it is the more reactive metal / it supplies electrons / it forms ions more readily than iron [1]

(d) replace zinc with magnesium
replace iron with copper
use (more) concentrated sulfuric acid
accept use a more concentrated acid / a more concentrated solution

any **two** [2]

3 A student used the following electrochemical cell.

The reading on the voltmeter was +1.10 V.



(i) Draw an arrow on the diagram to show the direction of electron flow. [1]

(ii) Suggest the change, if any, in the voltmeter reading if the zinc electrode was replaced with an iron electrode. Explain your answer.

.....

..... [2]

(iii) The zinc electrode was replaced with a silver electrode. The reading on the voltmeter was –0.46 V.

Suggest why the sign of the voltmeter reading became negative.

.....

..... [1]

MARKING SCHEME:

(i)	arrow (anywhere) going from Zn → Cu	**1**
(ii)	reading would decrease (1) Fe less reactive than Zn (1) **OR** difference in reactivity (between Fe and Cu) is smaller	**2**
(iii)	Ag less reactive than Cu	**1**

FUEL CELLS

1 The Periodic Table is very useful to chemists.

Refer only to elements with atomic numbers 1 to 36 in the Periodic Table provided when answering **Question 1**.

One element in the first 36 elements is used as the fuel in a fuel cell.

(i) Name this element.

..... [1]

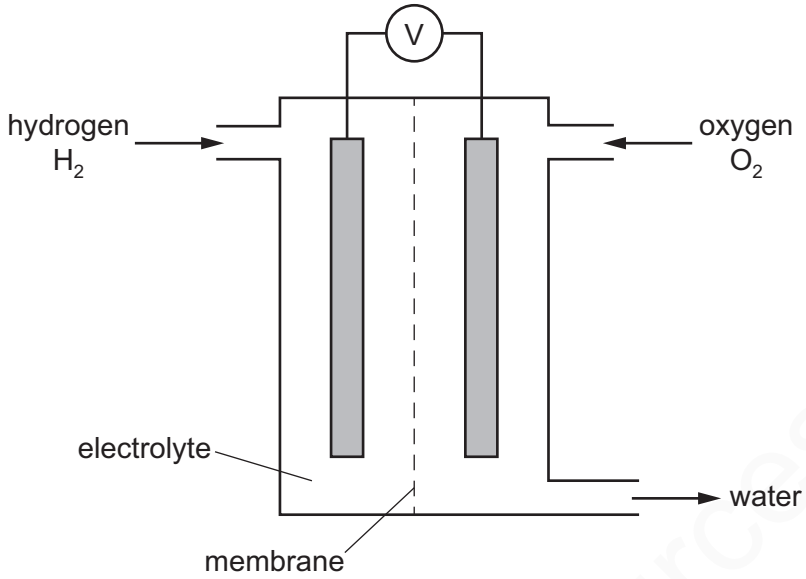
(ii) Write the overall chemical equation for the reaction which occurs when the element in **(c)(i)** reacts in a fuel cell.

..... [2]

MARKING SCHEME:

(i)	gydrogen / H	1
(ii)	$2H_2 + O_2 \rightarrow 2H_2O$ water as product from reaction of hydrogen and oxygen (1) balanced (1)	2

2 Hydrogen and oxygen react together in a hydrogen fuel cell. A hydrogen fuel cell is shown in the diagram.



(a) Name the process by which oxygen is obtained from air.

..... [1]

(b) **(i)** In a hydrogen fuel cell, the hydrogen molecules are converted into hydrogen ions, H^+, according to the ionic half-equation shown.

$$H_2 \rightarrow 2H^+ + 2e^-$$

What type of reaction does this ionic half-equation represent?

..... [1]

(ii) What **type** of substance reacts by donating hydrogen ions, H^+?

..... [1]

(c) Write a chemical equation for the overall reaction that occurs in a hydrogen fuel cell.

..... [1]

(d) Hydrogen fuel cells are being developed as alternatives to petrol engines in cars.

(i) Give **one** advantage of hydrogen fuel cells compared to petrol engines.

..... [1]

(ii) Give **one** disadvantage of hydrogen fuel cells compared to petrol engines.

..... [1]

(e) Some fuel cells use ethanol, C_2H_5OH, instead of hydrogen. Carbon dioxide and water are products of the reaction in an ethanol fuel cell.

(i) Write a chemical equation for the overall reaction occurring in an ethanol fuel cell.

..... [2]

(ii) State an environmental problem caused by the release of carbon dioxide into the atmosphere.

..... [1]

(iii) Name the process by which ethanol can be manufactured from a renewable resource.

..... [1]

(f) Name the process occurring when electrical energy is used to break down an ionic compound.

..... [1]

MARKING SCHEME:

(a)	fractional distillation	**1**
(b)(i)	oxidation	**1**
(b)(ii)	acid(ic)	**1**
(c)	$2H_2 + O_2 \rightarrow 2H_2O$	**1**
(d)(i)	no carbon dioxide produced / more efficient	**1**
(d)(ii)	storage of hydrogen is difficult / takes more space to store (hydrogen) / high likelihood of (hydrogen) leaks / lack of availability of hydrogen	**1**
(e)(i)	$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$ **M1** species correct **M2** balanced	**2**
(e)(ii)	climate change / greenhouse effect / consequence of climate change	**1**
(e)(iii)	fermentation	**1**
(f)	electrolysis	**1**

SMART EXAM RESOURCES
SUBJECT: CHEMISTRY
TOPIC:ELECTROCHEMISTRY
SET-2-QP-MS

FUEL CELLS

1 Hydrogen–oxygen fuel cells can be used to produce electricity in vehicles.

(i) Write the symbol equation for the overall reaction in a hydrogen–oxygen fuel cell.

..... [2]

(ii) State **one** advantage of using hydrogen–oxygen fuel cells instead of petrol in vehicle engines.

..... [1]

MARK SCHEME:

(i)	$2H_2 + O_2 \rightarrow 2H_2O$ **M1** all formulae(1) **M2** equation correct(1)	**2**
(ii)	no carbon dioxide evolved **OR** more efficient	**1**

2 Hydrogen–oxygen fuel cells can be used to produce electricity to power cars.
Petrol produces carbon dioxide and carbon monoxide when it powers cars.

(i) State **one** adverse effect of carbon dioxide and carbon monoxide.

carbon dioxide .....

carbon monoxide .....
[2]

(ii) State **one** disadvantage, other than cost, of using hydrogen–oxygen fuel cells to power cars compared to using petrol.

..... [1]

MARK SCHEME:

(i)	**M1** carbon dioxide: (increased) global warming **M2** carbon monoxide: toxic	**2**
(ii)	needs high pressure to store hydrogen	**1**