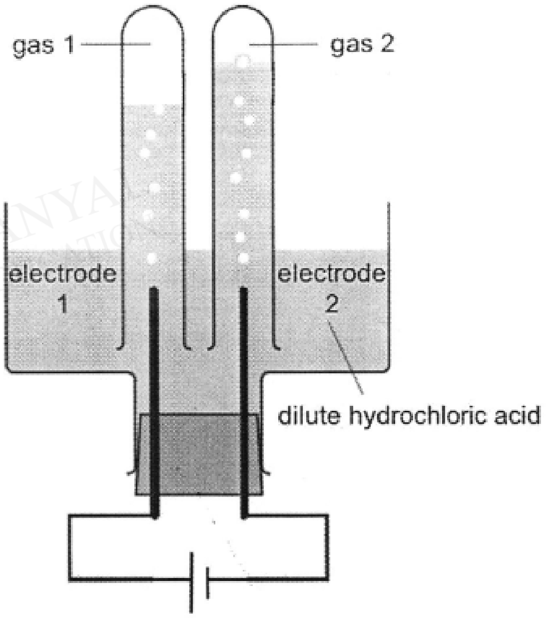
# O Level Pure Chemistry Structured

## Electrolysis Test 1.0

Q1

The diagram below shows a set-up used to electrolyse dilute hydrochloric acid.




**(i)** Write the half-equations for the reactions occurring at each electrode.

electrode 1:

electrode 2:

[2]

**(ii)** The actual volume of gas 2 collected is less than expected. Suggest a reason for this.

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..................................................................................................[1]

**(iii)** Describe and explain, with the aid of a half-equation, any difference you would expect to observe if the electrolyte was changed to concentrated hydrochloric acid.

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

[total = 11 marks]

Q2

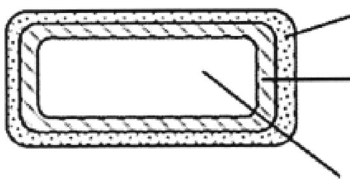
Galvanisation is the process of coating the entire surface of a piece of iron with zinc to prevent it from rusting. Two common ways of galvanising iron are hot-dip galvanisation and electro-galvanisation.

**(a) Hot-dip galvanisation**

The piece of iron to be galvanised is dipped into a molten bath of zinc at a temperature of around 460°C. The piece of iron is then cooled and exposed to the air. The outermost layer of zinc then reacts with oxygen and carbon dioxide in the air as follows:

reaction 1: zinc reacts with oxygen to form zinc oxide
reaction 2: zinc oxide reacts with carbon dioxide to form zinc carbonate

The resulting iron piece is as shown.




**(c) Electro-galvanisation (electroplating an object with zinc)**

The piece of iron to be galvanised and a piece of zinc are used as electrodes and dipped into an electrolyte containing a mixture of aqueous zinc cyanide, $Zn(CN)_2$, and aqueous sodium hydroxide at room temperature and pressure. An external electrical power supply is used. Zinc ions are discharged to form zinc atoms, which are coated onto the piece of iron.

**(i)** Draw a labelled diagram of the experimental setup for electro-galvanisation.

[2]

**(ii)** What is the formula for the cyanide ion?

..........................................................................................................[1]

**(iii)** Some processes of electro-galvanisation employ the use of dilute acids in the electrolyte instead of aqueous sodium hydroxide.

Explain what problem this could pose.

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..........................................................................................................[1]

**(d)** Suggest an **advantage** that electro-galvanisation has over hot-dip galvanisation.

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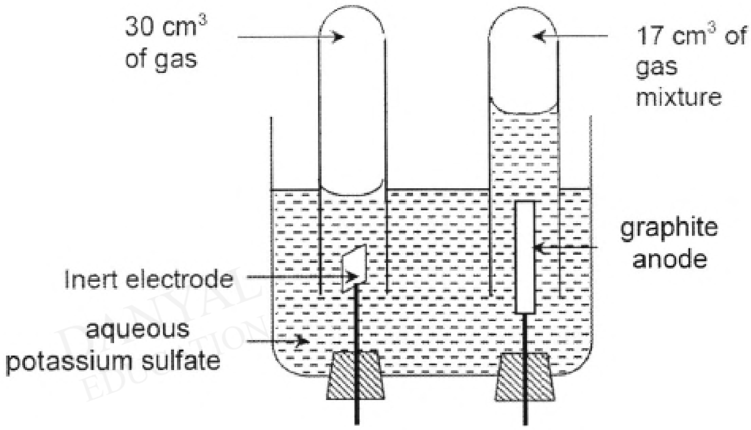
Q3

The diagram below shows the electrolysis of an aqueous solution of potassium sulfate using inert electrodes.




**(a)** Write equations for the reactions that happen at each electrode during the electrolysis of aqueous potassium sulfate. Include state symbols.

At the cathode : ..............................................................................................

At the anode : ..............................................................................................

[3]

**(b)** When graphite anode and a very high current are used in this electrolysis, the gas liberated is a mixture of oxygen, carbon monoxide and carbon dioxide. In the experiment illustrated above, 30 $cm^3$ of gas formed above the cathode and 17 $cm^3$ of gas formed above the anode.

**(i)** Explain, with the help of **two** equations, why the oxides of carbon are produced at the anode.

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..............................................................................................[3]

**(ii)** Using the equations in **b(i)**, explain why the volume of gas collected at the anode is larger than expected.

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..............................................................................................[1]

**(iii)** The gas at the anode was collected and its volume was reduced to 9 $cm^3$ when shaken with aqueous sodium hydroxide.
Deduce the volume of carbon dioxide in the gas mixture at the anode and explain the reaction that results in the reduction of volume.

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

**(c)** An experiment is set up to electroplate a fresh flower with silver.
Suggest why the fresh flower must be coated with carbon particles first.

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[Total: 10]

Q4

**A1** Magnesium is best known for burning with a characteristic brilliant white light. The metal itself was first produced by Sir Humphry Davy in 1808 by the electrolysis of a molten mixture of magnesia (MgO) and mercury oxide. Mercury oxide was added as an impurity to lower the melting point of magnesia, and inert electrodes were used during the electrolysis.

**(a)** **(i)** Davy obtained magnesium at the cathode.
Write an ionic equation to show how magnesium was formed at the cathode.
Suggest a reason why the magnesium was obtained as a mixture at the cathode.

ionic equation ..................................................

reason ..........................................................................................
[2]

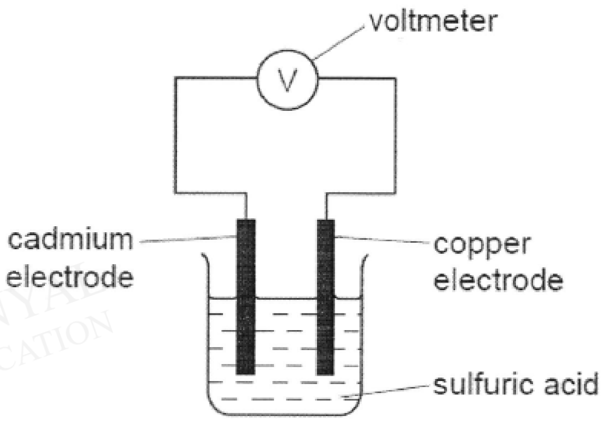
**(ii)** What product did Davy obtain at the anode?
Write an ionic equation for the reaction at the anode.

product ..................................................

ionic equation ..........................................................................................
[2]

Q5

**(a)** A reactivity series can 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.

| cell | electrode 1 (–) | electrode 2 (+) | voltage / volts |
|---|---|---|---|
| 1 | cadmium | copper | 0.74 |
| 2 | tin | copper | 0.48 |
| 3 | zinc | copper | 1.10 |

**(i)** What is a simple cell?

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

**(ii)** Place the four metals in order of increasing reactivity and explain how you used the data in the table to arrive at this order.

least reactive ..................................................

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most reactive ..................................................

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................................................................................................................................. [3]

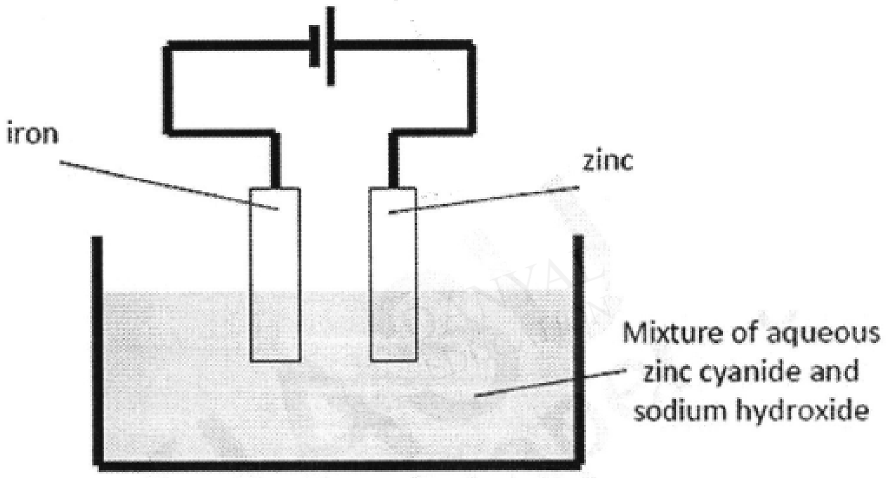
**Answers**

**Electrolysis Test 1.0**

Q1

| | | |
|---|---|---|
| (i) | electrode 1 (cathode): $2H^{+} (aq) + 2e^{-}$ □ $H_2 (g)$ [1]<br>electrode 2 (anode): $4OH^{-} (aq)$ □ $O_2 (g) + 2H_2O (l) + 4e^{-}$ [1] | **[2]** |
| (ii) | Oxygen gas is soluble in water and some dissolves. | **[1]** |
| (iii) | If concentrated hydrochloric acid is used, chloride ions would be discharged instead to produce chlorine gas, so the volume of gas collected at electrode 2 would be the same as at electrode 1 / a greenish-yellow gas is produced instead. [1]<br>$2Cl^{-} (aq)$ □ $Cl_2 (g) + 2e^{-}$ [1] | **[2]** |

Q2

<table>
<tr><td>c(i)</td><td>


Iron electrode connected to negative electrode of cell.</td><td>correct electrodes -1m<br><br>correct electrolyte -1m</td></tr>
<tr><td>(ii)</td><td>$CN^-$</td><td>1</td></tr>
<tr><td>(iii)</td><td>The acids in the electrolyte will react with the zinc and iron pieces.</td><td>1</td></tr>
<tr><td>d</td><td>Electro-galvanisation is carried out at room temperature and no heating is needed.<br>Hot-dip galvanisation is carried out at 460°C.</td><td>1</td></tr>
</table>

Q3

| | | |
|---|---|---|
| OR B9(a) | At the cathode : $2\ H^+\ (aq) + 2e^- \rightarrow H_2\ (g)$<br>At the anode : $4OH^-\ (aq) \rightarrow 2H_2O\ (l) + O_2\ (g) + 4e^-$ | correct equations – 1m each<br>correct state symbols – 1m |
| b(i) | Oxygen produced reacts with the carbon electrode to form carbon dioxide, which continues to react with the carbon electrode to form carbon monoxide.<br><br>$C + O_2 \rightarrow CO_2$<br>$C + CO_2 \rightarrow 2CO$ or<br>$2C + O_2 \rightarrow 2CO$ | 1<br><br><br><br>1<br>1 |
| (ii) | 1 mol of oxygen produces 2 mols carbon monoxide. | 1 |
| (iii) | 8 $cm^3$<br>Carbon dioxide, an acidic oxide reacts with the alkali, sodium hydroxide. | 1<br>1 |
| c | Carbon is a conductor of electricity. | 1 |

Q4

(a)(i) $Mg^{2+} + 2e \rightarrow Mg$ [1]

Mercury was obtained with the magnesium [1]

(ii) oxygen [1]

$2O^{2-} \rightarrow O_2 + 4e$ [1]

Q5

(a) (i) device which changes chemical energy [1] into electrical energy; [1]
OR
produces a voltage / potential difference / electricity [1] due to difference in reactivity of two metals; [1]
OR
produces a voltage / potential difference / electricity [1] by redox reactions [1]

(ii) 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 form the negative electrode or pole / copper is least reactive because it is the positive electrode or pole in all the cells / the bigger the difference in reactivity, the bigger the voltage / zinc has highest voltage because it is most reactive / more reactive metals give higher voltage