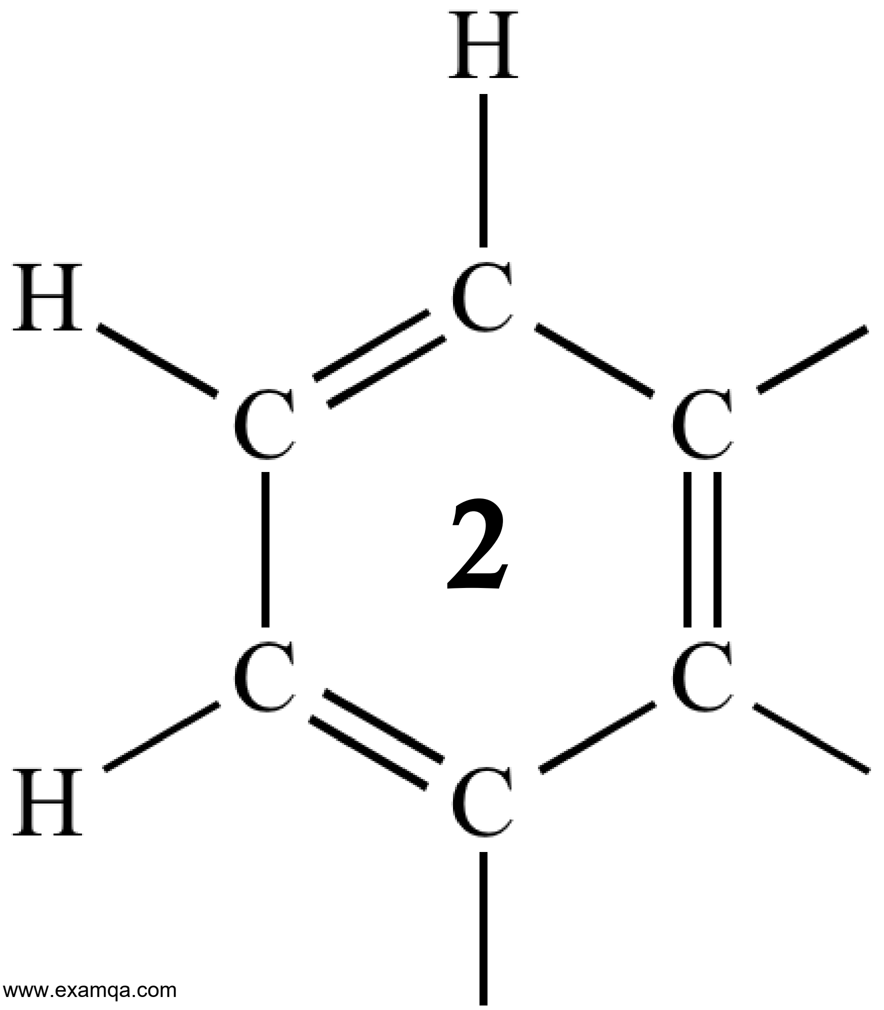
LONDON SCIENCE COLLEGE

OCR A2 CHEMISTRY

MODULE 5.5

REDOX

ELECTRODE POTENTIALS



1 In a test, aqueous iron(III) ions are reduced to aqueous iron(II) ions by iodide ions.
This reaction could be used to provide electrical energy in a cell.

(a) The standard electrode potential for the reduction of iron(III) ions into iron(II) ions can be measured by connecting a suitable electrode to a standard hydrogen electrode.
Draw a clearly labelled diagram to show the components and reagents, including their concentrations, in this Fe(III)/Fe(II) electrode.
Do **not** draw the salt bridge or the standard hydrogen electrode.

(3)

(b) A salt bridge is used to complete the cell. This could be prepared using potassium nitrate solution and filter paper.

State the purpose of the salt bridge. State **one** essential requirement of the soluble ionic compound used to make the salt bridge.

Purpose of salt bridge ....

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Requirement ....

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(2)
(Total 5 marks)

2 One cell that has been used to provide electrical energy is the Daniell cell. This cell uses copper and zinc.

(a) The conventional representation for the Daniell cell is

$$Zn(s) \mid Zn^{2+}(aq) \mid\mid Cu^{2+}(aq) \mid Cu(s)$$

The e.m.f. of this cell under standard conditions is +1.10 V.

Deduce the half-equations for the reactions occurring at the electrodes.

At Zn electrode ....

At Cu electrode ....

(2)

(b) A Daniell cell was set up using 100 cm^3 of a 1.0 mol dm^{-3} copper(II) sulfate solution. The cell was allowed to produce electricity until the concentration of the copper(II) ions had decreased to 0.50 mol dm^{-3}.

Calculate the decrease in mass of the zinc electrode. Show your working.

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(3)

(c) You are provided with the Daniell cell referred to in part (b), including a zinc electrode of known mass.

Briefly outline how you would carry out an experiment to confirm your answer to part (b).

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(3)
(Total 8 marks)

3 Copper, in the form of nanoparticles of copper(II) hexacyanoferrate(II), has recently been investigated as an efficient method of storing electrical energy in a rechargeable cell.

(a) Solar cells generate an electric current from sunlight. These cells are often used to provide electrical energy for illuminated road signs.

Explain why rechargeable cells are connected to these solar cells.

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(2)

(b) Suggest **one** reason why many waste disposal centres contain a separate section for cells and batteries.

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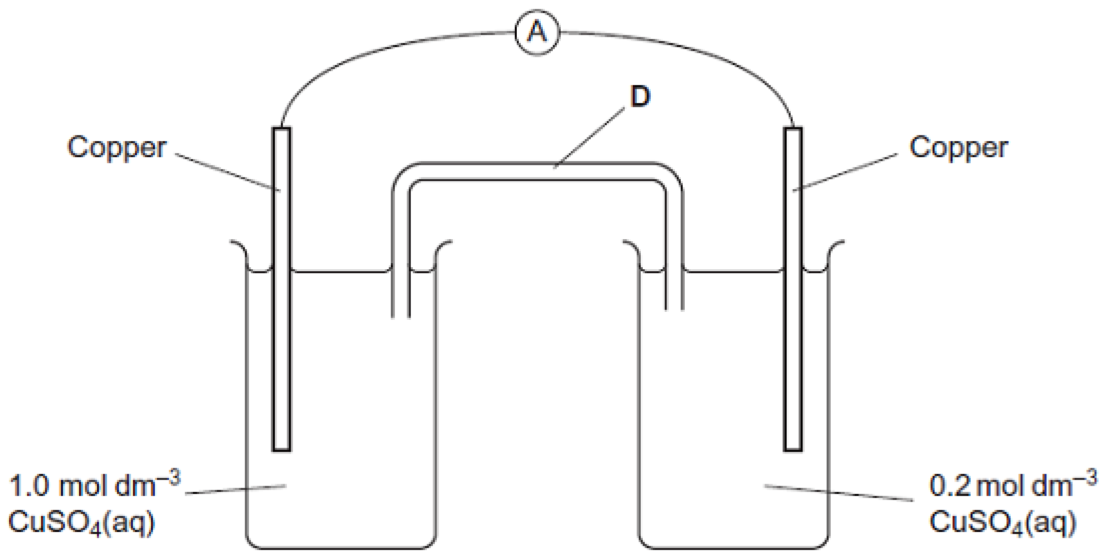
(1)
(Total 3 marks)

4 An electrochemical cell is shown in the diagram. In this cell, the amount of copper in the electrodes is much greater than the amount of copper ions in the copper sulfate solutions.



(a) Explain how the salt bridge **D** provides an electrical connection between the two electrodes.

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(1)

(b) Suggest why potassium chloride would **not** be a suitable salt for the salt bridge in this cell.

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(1)

(c) In the external circuit of this cell, the electrons flow through the ammeter from right to left.

Suggest why the electrons move in this direction.

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(2)

(d) Explain why the current in the external circuit of this cell falls to zero after the cell has operated for some time.

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(1)

(e) The simplified electrode reactions in a rechargeable lithium cell are

Electrode **A** $Li^+ + MnO_2 + e^- \longrightarrow LiMnO_2$ $E = -0.15$ V

Electrode **B** $Li^+ + e^- \longrightarrow Li$

Electrode **B** is the negative electrode.

(i) The e.m.f. of this cell is 2.90 V.

Use this information to calculate a value for the electrode potential of electrode **B**.

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(1)

(ii) Write an equation for the overall reaction that occurs when this lithium cell is being **recharged**.

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(2)

(iii) Suggest why the recharging of a lithium cell may lead to release of carbon dioxide into the atmosphere.

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(1)

(Total 9 marks)

5 An experiment was carried out to measure the e.m.f. of this cell.

$$Al(s) \mid Al^{3+}(aq) \parallel H^+(aq) \mid H_2(g) \mid Pt(s)$$

(a) The aluminium used as the electrode is rubbed with sandpaper prior to use.

Suggest the reason for this.

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(1)

(b) Draw a labelled diagram of a suitable apparatus for the right-hand electrode in this cell. You do **not** need to include the salt bridge or the external electrical circuit.

(2)

(c) A simple salt bridge can be prepared by dipping a piece of filter paper into potassium carbonate solution. Explain why such a salt bridge would **not** be suitable for use in this cell.

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(2)
(Total 5 marks)

6 This table shows some standard electrode potential data.

Electrode half-equation	**$E^\ominus$ / V**
$Au^+(aq) + e^- \longrightarrow Au(s)$	+1.68
$\frac{1}{2}O_2(g) + 2H^+(aq) + 2e^- \longrightarrow H_2O(l)$	+1.23
$Ag^+(aq) + e^- \longrightarrow Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq)$	+0.77
$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$	+0.34
$Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$	−0.44

(a) Draw a labelled diagram of the apparatus that could be connected to a standard hydrogen electrode in order to measure the standard electrode potential of the Fe^{3+} / Fe^{2+} electrode.

In your diagram, show how this electrode is connected to the standard hydrogen electrode and to a voltmeter. Do **not** draw the standard hydrogen electrode.

State the conditions under which this cell should be operated in order to measure the standard electrode potential.

Conditions ....

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(5)

(b) Use data from the table to deduce the equation for the overall cell reaction of a cell that has an e.m.f. of 0.78 V.
Give the conventional cell representation for this cell.
Identify the positive electrode.

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(4)

(c) Use data from the table to explain why Au^+ ions are **not** normally found in aqueous solution.

Write an equation to show how Au^+ ions would react with water.

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(3)

(d) Use data from the table to predict and explain the redox reactions that occur when iron powder is added to an excess of aqueous silver nitrate.

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(3)
(Total 15 marks)

7 The table shows some electrode half-equations and the associated standard electrode potentials.

Equation number	Electrode half-equation	$E^\ominus$ / V
1	$Cd(OH)_2(s) + 2e^- \rightarrow Cd(s) + 2OH^-(aq)$	–0.88
2	$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	–0.76
3	$NiO(OH)(s) + H_2O(l) + e^- \rightarrow Ni(OH)_2(s) + OH^-(aq)$	+0.52
4	$MnO_2(s) + H_2O(l) + e^- \rightarrow MnO(OH)(s) + OH^-(aq)$	+0.74
5	$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$	+1.23

(a) In terms of electrons, state the meaning of the term *oxidising agent*.

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(1)

(b) Deduce the identity of the weakest oxidising agent in the table.
Explain how $E^\ominus$ values can be used to make this deduction.

Weakest oxidising agent ....

Explanation ....

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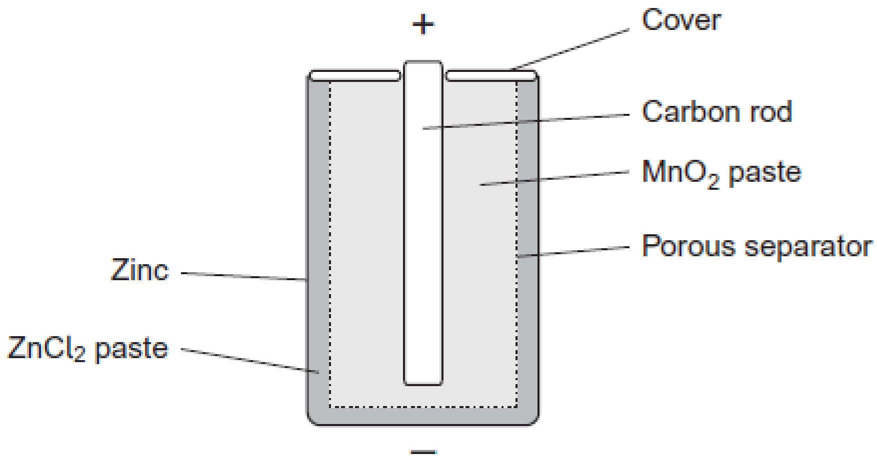
(2)

(c) The diagram shows a non-rechargeable cell that can be used to power electronic devices. The relevant half-equations for this cell are equations **2** and **4** in the table above.



(i) Calculate the e.m.f. of this cell.

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(1)

(ii) Write an equation for the overall reaction that occurs when the cell discharges.

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(1)

(iii) Deduce **one** essential property of the non-reactive porous separator labelled in the diagram.

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(1)

(iv) Suggest the function of the carbon rod in the cell.

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(1)

(v) The zinc electrode acts as a container for the cell and is protected from external damage. Suggest why a cell often leaks after being used for a long time.

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(1)

(d) A rechargeable nickel–cadmium cell is an alternative to the cell shown in part (c). The relevant half-equations for this cell are equations **1** and **3** in the table above.

(i) Deduce the oxidation state of the nickel in this cell after recharging is complete. Write an equation for the overall reaction that occurs when the cell is **recharged**.

Oxidation state ....

Equation ....

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(3)

(ii) State **one** environmental advantage of this rechargeable cell compared with the non-rechargeable cell described in part (c).

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(1)

(e) An ethanol–oxygen fuel cell may be an alternative to a hydrogen–oxygen fuel cell. When the cell operates, all of the carbon atoms in the ethanol molecules are converted into carbon dioxide.

(i) Deduce the equation for the overall reaction that occurs in the ethanol–oxygen fuel cell.

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(1)

(ii) Deduce a half-equation for the reaction at the ethanol electrode. In this half-equation, ethanol reacts with water to form carbon dioxide and hydrogen ions.

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(1)

(iii) The e.m.f. of an ethanol–oxygen fuel cell is 1.00 V. Use data from the table above to calculate a value for the electrode potential of the ethanol electrode.

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(1)

(iv) Suggest why ethanol can be considered to be a carbon-neutral fuel.

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(2)
(Total 17 marks)