Q1. 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)

Q2. 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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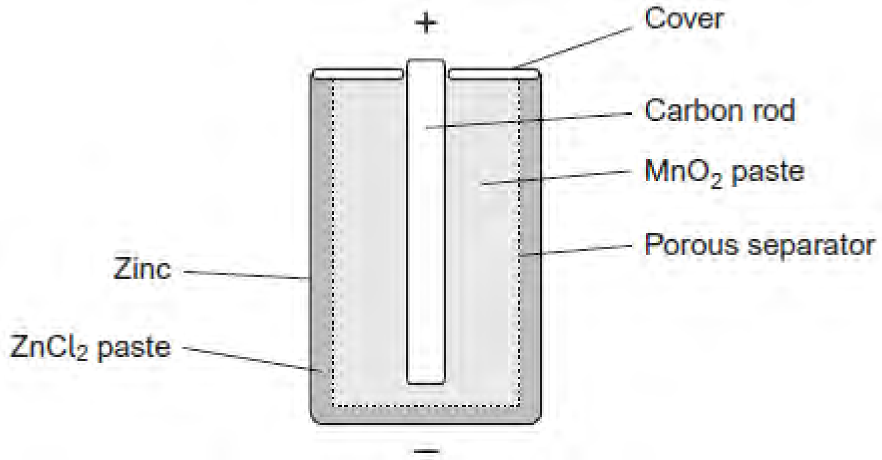
(2)

(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 ....

....

.... **(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.

.... **(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)

Q3. Redox reactions occur in the discharge of all electrochemical cells. Some of these cells are of commercial value.
The table below shows some redox half-equations and standard electrode potentials.

Half-equation	$E^{\ominus}$/ V
$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s)$	–0.76

$Ag_2O(s) + 2H^+(aq) + 2e– \longrightarrow 2Ag(s) + H_2O(l)$	+0.34
$O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l)$	+1.23
$F_2(g) + 2e^- \longrightarrow 2F^-(aq)$	+2.87

(a) In terms of electrons, state what happens to a reducing agent in a redox reaction.

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(1)

(b) Use the table above to identify the strongest reducing agent from the species in the table.

Explain how you deduced your answer.

Strongest reducing agent ....

Explanation ....

....

(2)

(c) Use data from the table to explain why fluorine reacts with water.

Write an equation for the reaction that occurs.

Explanation ....

....

....

Equation ....

....

(3)

(d) An electrochemical cell can be constructed using a zinc electrode and an electrode in which silver is in contact with silver oxide. This cell can be used to power electronic devices.

(i) Give the conventional representation for this cell.

....

(2)

(ii) Calculate the e.m.f. of the cell.

....
(1)

(iii) Suggest **one** reason why the cell cannot be electrically recharged.

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(1)

(e) The electrode half-equations in a lead–acid cell are shown in the table below.

Half-equation	$E^\ominus$/ V
$PbO_2(s) + 3H^+(aq) + HSO_4^-(aq) + 2e^- \longrightarrow PbSO_4(s) + 2H_2O(l)$	+1.69
$PbSO_4(s) + H^+(aq) + 2e^- \longrightarrow Pb(s) + HSO_4^-(aq)$	to be calculated

(i) The PbO_2/$PbSO_4$ electrode is the positive terminal of the cell and the e.m.f. of the cell is 2.15 V.

Use this information to calculate the missing electrode potential for the half-equation shown in the table.

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(1)

(ii) A lead–acid cell can be recharged.
Write an equation for the overall reaction that occurs when the cell is being recharged.

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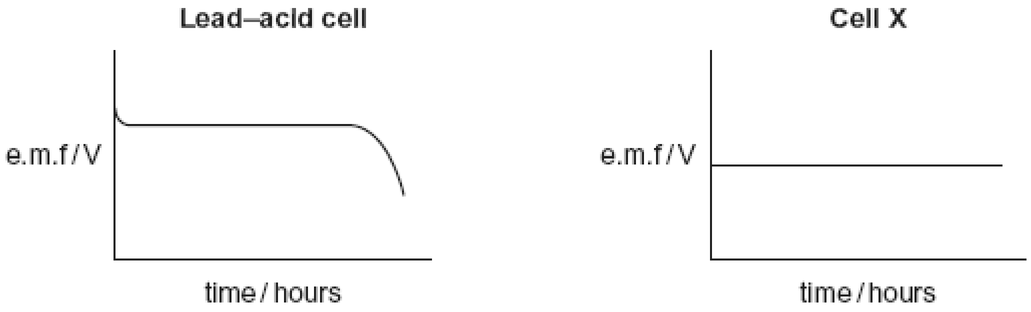
(2)

(f) The diagrams below show how the e.m.f. of each of two cells changes with time when each cell is used to provide an electric current.



(i) Give **one** reason why the e.m.f. of the **lead–acid cell** changes after several hours.

....

....

(1)

(ii) Identify the type of cell that behaves like **cell X**.

....

(1)

(iii) Explain why the voltage remains constant in **cell X**.

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(Extra space) ....

....

(2)
(Total 17 marks)

Q4. Nickel–cadmium cells are used to power electrical equipment such as drills and

shavers.
The electrode reactions are shown below.

$$NiO(OH) + H_2O + e^- \rightarrow Ni(OH)_2 + OH^- \qquad E^\ominus = +0.52\ V$$

$$Cd(OH)_2 + 2e^- \rightarrow Cd + 2OH^- \qquad E^\ominus = -0.88\ V$$

(a) Calculate the e.m.f. of a nickel–cadmium cell.

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(1)

(b) Deduce an overall equation for the reaction that occurs in the cell when it is used.

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....

....

(2)

(c) Identify the oxidising agent in the overall cell reaction and give the oxidation state of the metal in this oxidising agent.

Oxidising agent ....

Oxidation state ....

(2)
(Total 5 marks)