

Q1. This table shows some standard electrode potential data.

Electrode half-equation	E^\ominus / V
$\text{Au}^+(\text{aq}) + \text{e}^- \longrightarrow \text{Au}(\text{s})$	+1.68
$\frac{1}{2}\text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{H}_2\text{O}(\text{l})$	+1.23
$\text{Ag}^+(\text{aq}) + \text{e}^- \longrightarrow \text{Ag}(\text{s})$	+0.80
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \longrightarrow \text{Fe}^{2+}(\text{aq})$	+0.77
$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Cu}(\text{s})$	+0.34
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Fe}(\text{s})$	-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 $\text{Fe}^{3+} / \text{Fe}^{2+}$ 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	$\text{Cd}(\text{OH})_2(\text{s}) + 2\text{e}^- \rightarrow \text{Cd}(\text{s}) + 2\text{OH}^-(\text{aq})$	-0.88
2	$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn}(\text{s})$	-0.76
3	$\text{NiO}(\text{OH})(\text{s}) + \text{H}_2\text{O}(\text{l}) + \text{e}^- \rightarrow \text{Ni}(\text{OH})_2(\text{s}) + \text{OH}^-(\text{aq})$	+0.52
4	$\text{MnO}_2(\text{s}) + \text{H}_2\text{O}(\text{l}) + \text{e}^- \rightarrow \text{MnO}(\text{OH})(\text{s}) + \text{OH}^-(\text{aq})$	+0.74
5	$\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l})$	+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.

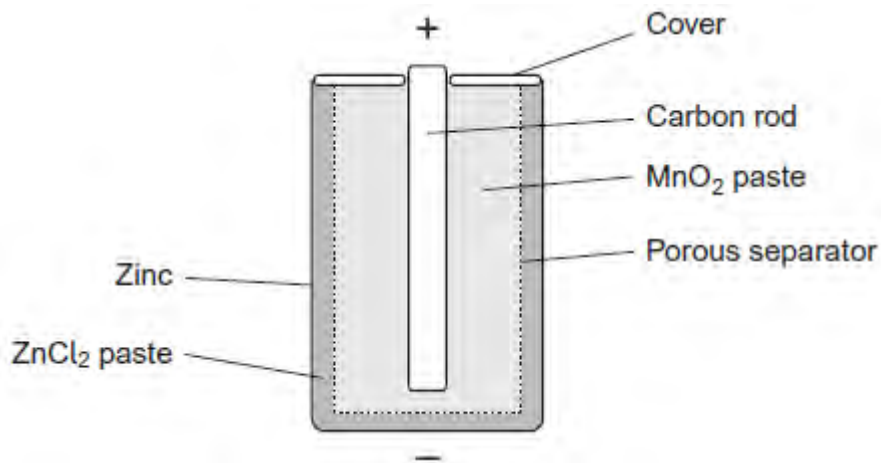
Weakest oxidising agent

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)

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
$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn}(\text{s})$	-0.76

$\text{Ag}_2\text{O}(\text{s}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow 2\text{Ag}(\text{s}) + \text{H}_2\text{O}(\text{l})$	+0.34
$\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l})$	+1.23
$\text{F}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{F}^-(\text{aq})$	+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

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(2)

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

Write an equation for the reaction that occurs.

Explanation

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Equation

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(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.

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(2)

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

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(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
$\text{PbO}_2(\text{s}) + 3\text{H}^+(\text{aq}) + \text{HSO}_4^-(\text{aq}) + 2\text{e}^- \rightarrow \text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l})$	+1.69
$\text{PbSO}_4(\text{s}) + \text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{Pb}(\text{s}) + \text{HSO}_4^-(\text{aq})$	to be calculated

(i) The $\text{PbO}_2/\text{PbSO}_4$ 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.

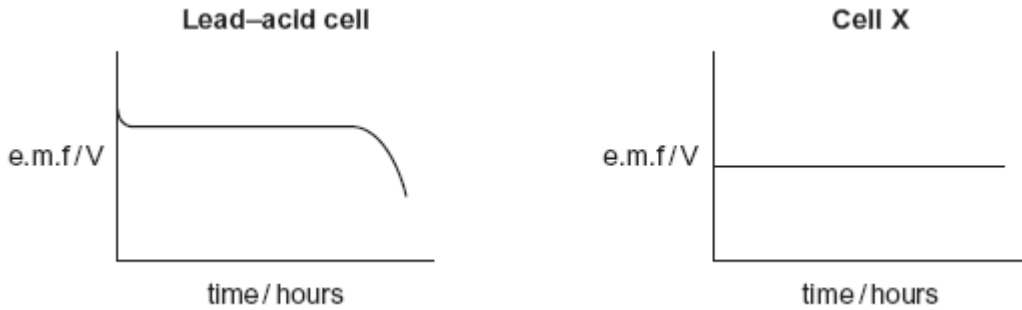
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(f) The diagrams below show how the e.m.f. of each of two cells changes with time when each cell is used to provide an electric current.



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

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(1)

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

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(1)

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

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(Extra space)

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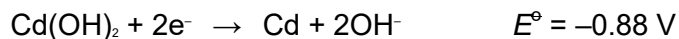
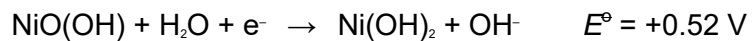
(2)

(Total 17 marks)

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

shavers.

The electrode reactions are shown below.



- (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)