Q1. Fuel cells are an increasingly important energy source for vehicles. Standard electrode potentials are used in understanding some familiar chemical reactions including those in fuel cells.

The following table contains some standard electrode potential data.

Electrode half-equation	$E^{\ominus}$ / V
$F_2 + 2e^- \longrightarrow 2F^-$	+2.87
$Cl_2 + 2e^- \longrightarrow 2Cl^-$	+1.36
$O_2 + 4H^+ + 4e^- \longrightarrow 2H_2O$	+1.23
$Br_2 + 2e^- \longrightarrow 2Br^-$	+1.07
$I_2 + 2e^- \longrightarrow 2I^-$	+0.54
$O_2 + 2H_2O + 4e^- \longrightarrow 4OH^-$	+0.40
$SO_4^{2-} + 4H^+ + 2e^- \longrightarrow SO_2 + 2H_2O$	+0.17
$2H^+ + 2e^- \longrightarrow H_2$	0.00
$4H_2O + 4e^- \longrightarrow 4OH^- + 2H_2$	−0.83

(a) A salt bridge was used in a cell to measure electrode potential.

Explain the function of the salt bridge.

.....

.....

.....

.....

(2)

(b) Use data from the table above to deduce the halide ion that is the weakest reducing agent.

.....

(1)

(c) Use data from the table to justify why sulfate ions should **not** be capable of oxidising bromide ions.

.....

.....

.....

(1)

(d) Use data from the table to calculate a value for the EMF of a hydrogen–oxygen fuel cell operating under alkaline conditions.

EMF = ... V

(1)

(e) There are two ways to use hydrogen as a fuel for cars. One way is in a fuel cell to power an electric motor, the other is as a fuel in an internal combustion engine.

Suggest the major advantage of using the fuel cell.

.....

.....

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(1)
(Total 6 marks)

Q2. 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.

.....

.....

.....

(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)

Q3. The table below shows some standard electrode potential data.

	$E^{\ominus}$ / V
$ZnO(s) + H_2O(l) + 2e^- \longrightarrow Zn(s) + 2OH^-(aq)$	–1.25
$Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$	–0.44

$O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$	+0.40
$2HOCl(aq) + 2H^+(aq) + 2e^- \longrightarrow Cl_2(g) + 2H_2O(l)$	+1.64

(a) Give the conventional representation of the cell that is used to measure the standard electrode potential of iron as shown in the table.

.....

(2)

(b) With reference to electrons, give the meaning of the term **reducing agent**.

.....

.....

(1)

(c) Identify the weakest reducing agent from the species in the table.

Explain how you deduced your answer.

Species.....

Explanation.....

.....

(2)

(d) When HOCl acts as an oxidising agent, one of the atoms in the molecule is reduced.

(i) Place a tick (✓) next to the atom that is reduced.

Atom that is reduced	**Tick (✓)**
H	
O	
Cl	

(1)

(ii) Explain your answer to part (i) in terms of the change in the oxidation state of this atom.

.....

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(1)

(e) Using the information given in the table, deduce an equation for the redox reaction that would occur when hydroxide ions are added to HOCl

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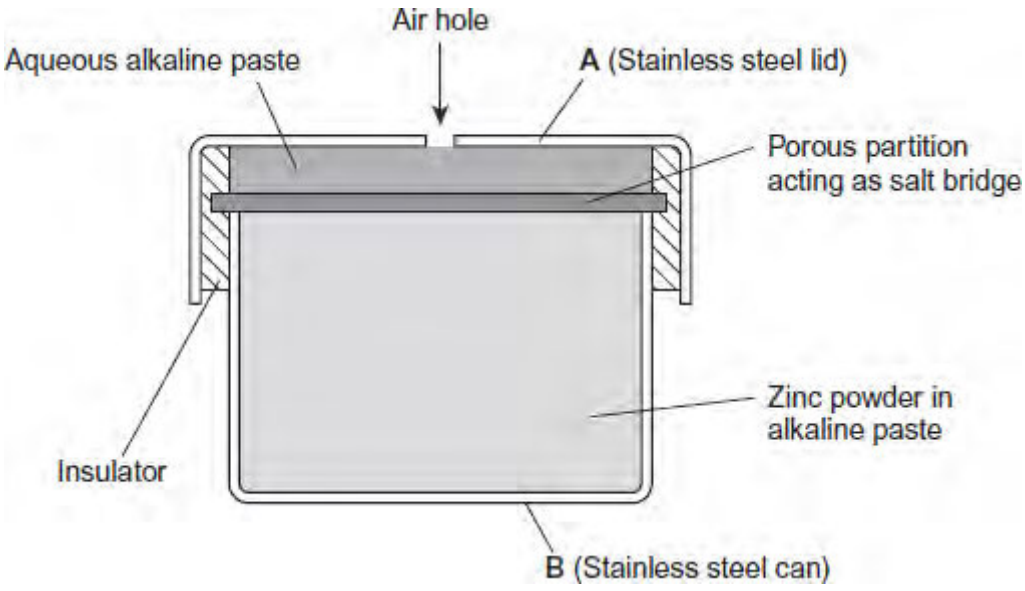
(2)

(f) The table is repeated to help you answer this question.

	$E^{\ominus}$ / V
$ZnO(s) + H_2O(l) + 2e^- \longrightarrow Zn(s) + 2OH^-(aq)$	–1.25
$Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$	–0.44
$O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$	+0.40
$2HOCl(aq) + 2H^+(aq) + 2e^- \longrightarrow Cl_2(g) + 2H_2O(l)$	+1.64

The half-equations from the table that involve zinc and oxygen are simplified versions of those that occur in hearing aid cells.

A simplified diagram of a hearing aid cell is shown in the following figure.



(i) Use data from the table to calculate the e.m.f. of this cell.

.....

.....

Answer = ...

(1)

(ii) Use half-equations from the table to construct an overall equation for the cell reaction.

.....

.....

.....

(1)

(iii) Identify which of **A** or **B**, in the figure, is the positive electrode. Give a reason for your answer.

Positive electrode .....

Reason .....

.....

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(2)

(iv) Suggest **one** reason, other than cost, why this type of cell is **not** recharged.

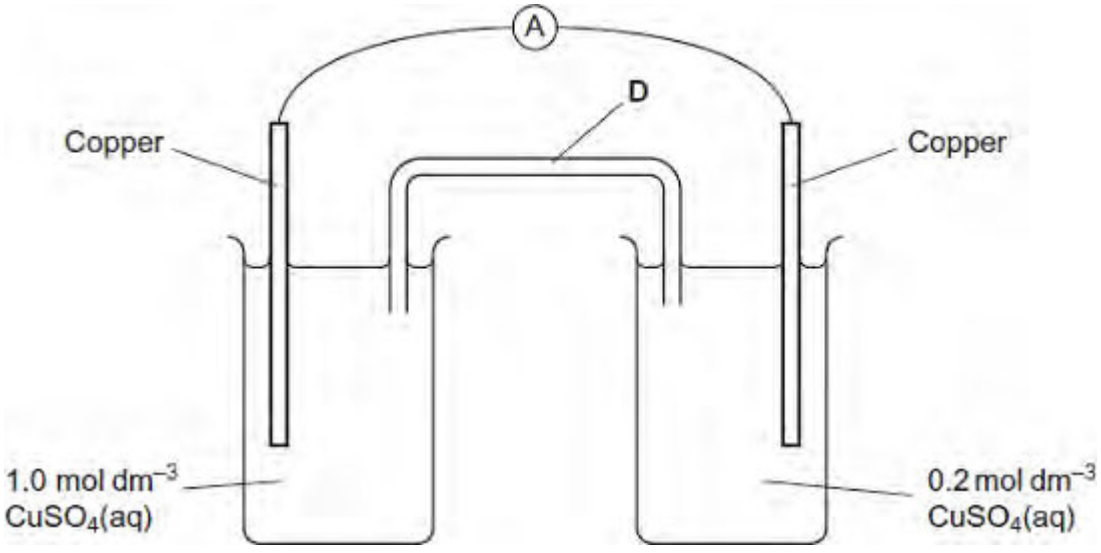
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(1)

(Total 14 marks)

Q4. 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.

.....

.....

(1)

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

.....

.....

(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.

.....

.....

(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**.

.....

.....

(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.

.....

.....

(1)

(Total 9 marks)

Q5. Hydrogen–oxygen fuel cells are used to provide electrical energy for electric motors in vehicles.

(a) In a hydrogen–oxygen fuel cell, a current is generated that can be used to drive an electric motor.

(i) Deduce half-equations for the electrode reactions in a hydrogen–oxygen fuel cell.

Half-equation 1 .....

Half-equation 2 .....

(2)

(ii) Use these half-equations to explain how an electric current can be generated.

.....

.....

.....

.....

(2)

(b) Explain why a fuel cell does **not** need to be recharged.

.....

.....

.....

(1)

(c) To provide energy for a vehicle, hydrogen can be used either in a fuel cell or in an internal combustion engine.

Suggest the main advantage of using hydrogen in a fuel cell rather than in an internal combustion engine.

.....

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(1)

(d) Identify **one** major hazard associated with the use of a hydrogen–oxygen fuel cell in a vehicle.

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(1)
(Total 7 marks)

Q6. 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)

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 .....

.....

.....

(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*.

.....

.....

(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.

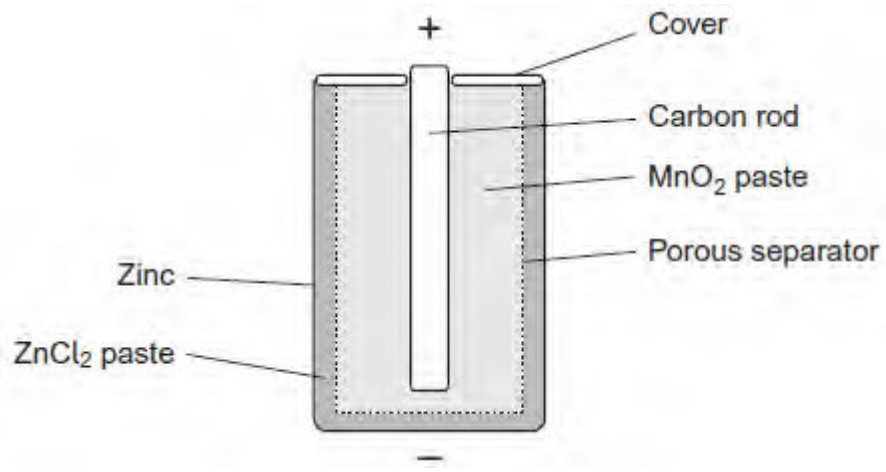
Weakest oxidising agent .....

Explanation .....

.....

(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.

.....

.....

.....

(1)

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

.....

.....

.....

(1)

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

.....

.....

(1)

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

.....

.....

(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.

.....

.....

(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).

.....

.....

(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.

.....

(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.

.....

.....

(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.

.....

(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.

.....

.....

(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.

.....

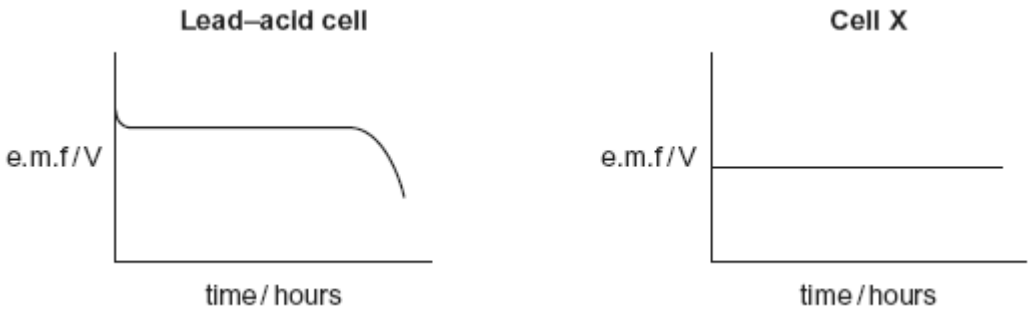
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(2)

(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.

.....

(1)

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

.....

.....

.....

(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)

Q1. (a) Lithium ion cells are used to power cameras and mobile phones.
A simplified representation of a cell is shown below.

$$Li \mid Li^+ \parallel Li^+, CoO_2 \mid LiCoO_2 \mid Pt$$

The reagents in the cell are absorbed onto powdered graphite that acts as a support medium. The support medium allows the ions to react in the absence of a solvent such as water.

The half-equation for the reaction at the positive electrode can be represented as follows.

$$Li^+ + CoO_2 + e^- \longrightarrow Li^+[CoO_2]^-$$

(i) Identify the element that undergoes a change in oxidation state at the positive electrode and deduce these oxidation states of the element.

Element .....

Oxidation state 1 .....

Oxidation state 2 .....

.....

(3)

(ii) Write a half-equation for the reaction at the negative electrode during operation of the lithium ion cell.

.....

(1)

(iii) Suggest two properties of platinum that make it suitable for use as an external electrical contact in the cell.

Property 1 .....

Property 2 .....

(2)

(iv) Suggest **one** reason why water is **not** used as a solvent in this cell.

.....

.....

(1)

(b) The half-equations for two electrodes used to make an electrochemical cell are shown below.

$ClO_3^-(aq) + 6H^+(aq) + 6e^- \longrightarrow Cl^-(aq) + 3H_2O(l)$ $\quad E^\ominus = +1.45$ V

$SO_4^{2-}(aq) + 2H^+(aq) + 2e^- \longrightarrow SO_3^{2-}(aq) + H_2O(l)$ $\quad E^\ominus = +0.17$ V

(i) Write the conventional representation for the cell using platinum contacts.

.....

(2)

(ii) Write an overall equation for the cell reaction and identify the oxidising and reducing agents.

Overall equation .....

.....

.....

Oxidising agent .....

Reducing agent .....

(3)

(Total 12 marks)

Q2. The electrons transferred in redox reactions can be used by electrochemical cells to provide energy.

Some electrode half-equations and their standard electrode potentials are shown in the table below.

Half-equation	$E^\ominus$**/V**
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$	+1.33
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.77
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$	–0.44

$Li^+(aq) + e^- \rightarrow Li(s)$	–3.04

(a) Describe a standard hydrogen electrode.

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(4)

(b) A conventional representation of a lithium cell is given below.
This cell has an e.m.f. of +2.91 V

$$Li(s) \mid Li^+(aq) \parallel Li^+(aq) \mid MnO_2(s) , LiMnO_2(s) \mid Pt(s)$$

Write a half-equation for the reaction that occurs at the positive electrode of this cell.

Calculate the standard electrode potential of this positive electrode.

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(2)

(c) Suggest what reactions occur, if any, when hydrogen gas is bubbled into a solution containing a mixture of iron(II) and iron(III) ions. Explain your answer.

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(2)

(d) A solution of iron(II) sulfate was prepared by dissolving 10.00 g of $FeSO_4.7H_2O$ (M_r = 277.9) in water and making up to 250 cm^3 of solution. The solution was left to stand, exposed to air, and some of the iron(II) ions became oxidised to iron(III) ions. A 25.0 cm^3 sample of the partially oxidised solution required 23.70 cm^3 of 0.0100 mol dm^{-3} potassium dichromate(VI) solution for complete reaction in the presence of an excess of dilute sulfuric acid.

Calculate the percentage of iron(II) ions that had been oxidised by the air.

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(6)
(Total 14 marks)

Q3. Hydrogen–oxygen fuel cells can operate in acidic or in alkaline conditions but commercial cells use porous platinum electrodes in contact with concentrated aqueous potassium hydroxide. The table below shows some standard electrode potentials measured in acidic and in alkaline conditions.

Half-equation	$E^{\ominus}$ /V
$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$	+1.23
$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$	+0.40
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$2H_2O(l) + 2e^- \rightarrow 2OH^-(aq) + H_2(g)$	– 0.83

(a) State why the electrode potential for the standard hydrogen electrode is equal to 0.00V.

.....

(1)

(b) Use data from the table to calculate the e.m.f. of a hydrogen–oxygen fuel cell operating in alkaline conditions.

.....

(1)

(c) Write the conventional representation for an alkaline hydrogen–oxygen fuel cell.

.....

(2)

(d) Use the appropriate half-equations to construct an overall equation for the reaction that occurs when an alkaline hydrogen–oxygen fuel cell operates. Show your working.

.....

.....

.....

.....

(2)

(e) Give **one** reason, other than cost, why the platinum electrodes are made by coating a porous ceramic material with platinum rather than by using platinum rods.

.....

.....

(1)

(f) Suggest why the e.m.f. of a hydrogen–oxygen fuel cell, operating in acidic conditions, is exactly the same as that of an alkaline fuel cell.

.....

(1)

(g) Other than its lack of pollution, state briefly the main advantage of a fuel cell over a re-chargeable cell such as the nickel–cadmium cell when used to provide power for an electric motor that propels a vehicle.

.....

.....

(1)

(h) Hydrogen–oxygen fuel cells are sometimes regarded as a source of energy that is carbon neutral. Give **one** reason why this may **not** be true.

.....

(1)
(Total 10 marks)

Q4. Use the standard electrode potential data in the table below to answer the questions which follow.

			$E^{\ominus}$ / V
$Ce^{4+}(aq) + e^-$	⇌	$Ce^{3+}(aq)$	+1.70
$MnO^-(aq) + 8H^+(aq) + 5e^-$	⇌	$Mn^{2+}(aq) + 4H_2O(l)$	+1.51
$Cl_2(g) + 2e^-$	⇌	$2Cl^-(aq)$	+1.36
$VO_2^+(aq) + 2H^+(aq) + e^-$	⇌	$VO^{2+}(aq) + H_2O(l)$	+1.00
$Fe^{3+}(aq) + e^-$	⇌	$Fe^{2+}(aq)$	+0.77

$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons H_2SO_3(aq) + H_2O(l)$ +0.17

(a) Name the standard reference electrode against which all other electrode potentials are measured.

.....

(1)

(b) When the standard electrode potential for $Fe^{3+}(aq) / Fe^{2+}(aq)$ is measured, a platinum electrode is required.

(i) What is the function of the platinum electrode?

.....

(ii) What are the standard conditions which apply to $Fe^{3+}(aq)/Fe^{2+}(aq)$ when measuring this potential?

.....

.....

.....

(3)

(c) The cell represented below was set up under standard conditions.

$$Pt|H_2SO_3(aq), SO_4^{2-}(aq)||MnO_4^-(aq), Mn^{2+}(aq)|Pt$$

Calculate the e.m.f. of this cell and write an equation for the spontaneous cell reaction.

Cell e.m.f. ….....

Equation .....

.....

(3)

(d) (i) Which one of the species given in the table is the strongest oxidising agent?

.....

(ii) Which of the species in the table could convert $Fe^{2+}(aq)$ into $Fe^{3+}(aq)$ but could not convert $Mn^{2+}(aq)$ into $MnO_4^-(aq)$?

.....

(3)

(e) Use data from the table of standard electrode potentials to deduce the cell which would have a standard e.m.f. of 0.93 V. Represent this cell using the convention shown in part (c).

.....

(2)
(Total 12 marks)

Q1. The table contains some standard electrode potential data.

Electrode half-equation	$E^{\ominus}$ / **V**
$F_2 + 2e^- \longrightarrow 2F^-$	+2.87
$Au^+ + e^- \longrightarrow Au$	+1.68
$2HOCl + 2H^+ + 2e^- \longrightarrow Cl_2 + 2H_2O$	+1.64
$Cl_2 + 2e^- \longrightarrow 2Cl^-$	+1.36
$O_2 + 4H^+ + 4e^- \longrightarrow 2H_2O$	+1.23
$Ag^+ + e^- \longrightarrow Ag$	+0.80
$Fe^{3+} + e^- \longrightarrow Fe^{2+}$	+0.77
$2H^+ + 2e^- \longrightarrow H_2$	0.00
$Fe^{2+} + 2e^- \longrightarrow Fe$	−0.44

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

.....

.....

(1)

(b) Identify the weakest oxidising agent in the table.
Explain your choice.

Weakest oxidising agent .....

Explanation .....

.....

.....

(2)

(c) Write the conventional representation of the cell used to measure the standard electrode potential for the Ag^+ / Ag electrode.

State the conditions necessary when measuring this value.

Conventional representation .....

.....

Conditions .....

.....

.....

.....

(4)

(d) Use data from the table to explain, in terms of redox, what happens when a soluble gold(I) compound containing Au^+ ions is added to water.

State what you would observe.

Write an equation for the reaction that occurs.

Explanation .....

.....

.....

.....

Observation .....

.....

Equation .....

(4)

The table is repeated below to help you answer these questions.

Electrode half-equation	$E^\ominus$ / **V**
$F_2 + 2e^- \longrightarrow 2F^-$	+2.87
$Au^+ + e^- \longrightarrow Au$	+1.68
$2HOCl + 2H^+ + 2e^- \longrightarrow Cl_2 + 2H_2O$	+1.64
$Cl_2 + 2e^- \longrightarrow 2Cl^-$	+1.36
$O_2 + 4H^+ + 4e^- \longrightarrow 2H_2O$	+1.23
$Ag^+ + e^- \longrightarrow Ag$	+0.80

$Fe^{3+} + e^- \longrightarrow Fe^{2+}$	+0.77
$2H^+ + 2e^- \longrightarrow H_2$	0.00
$Fe^{2+} + 2e^- \longrightarrow Fe$	−0.44

(e) A cell is made by connecting Fe^{2+} / Fe and Ag^+ / Ag electrodes with a salt bridge.

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

.....

.....

Answer .....

(1)

(ii) Suggest why potassium chloride would **not** be suitable for use in the salt bridge of this cell.

.....

.....

.....

(1)

(f) Use data from the table to explain what happens when a solution of iron(II) chloride is exposed to the air.

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(2)

(Total 15 marks)

Q2. 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 .....

.....

Requirement .....

.....

(2)
(Total 5 marks)

Q3. A biocide is a chemical that kills bacteria. A biocide is added to prevent the growth of bacteria in the water used in vases of flowers. Household bleach contains aqueous chlorine and can be used as the biocide. The concentration of chlorine in vase water decreases with time. It was decided to investigate the rate of this decrease.

The following experimental method was used to determine the concentration of chlorine in vase water at different times.

- A sample of vase water was taken.
- An excess of potassium iodide solution was added to the sample.
- The chlorine in the sample oxidised the I^- ions to I_2

- The iodine was titrated with sodium thiosulfate ($Na_2S_2O_3$) solution.
- These steps were repeated using further samples taken from the vase water at hourly intervals.

(a) Suggest **two** reasons why the concentration of chlorine in the vase water decreases with time.

Reason 1 .....

.....

Reason 2 .....

.....

(2)

(b) Suggest why this sampling technique has no effect on the rate at which the concentration of chlorine in the vase water decreases.

.....

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(1)

(c) Why was it important to use an **excess** of potassium iodide solution?

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(1)

(d) Use the following standard electrode potential data to explain why I_2 oxidises $S_2O_3^{2-}$ under standard conditions.

$$\tfrac{1}{2}I_2 + e^- \longrightarrow I^- \qquad E^\ominus = +0.54\ V$$

$$\tfrac{1}{2}S_4O_6^{2-} + e^- \longrightarrow S_2O_3^{2-} \qquad E^\ominus = +0.09\ V$$

.....

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.....

(1)

(e) Deduce an ionic equation for the reaction between I_2 and $S_2O_3^{2-}$

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(1)

(Total 6 marks)

Q4. The table below shows some standard electrode potential data.

	$E^{\ominus}$ / V
$ZnO(s) + H_2O(l) + 2e^- \longrightarrow Zn(s) + 2OH^-(aq)$	–1.25
$Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$	–0.44
$O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$	+0.40
$2HOCl(aq) + 2H^+(aq) + 2e^- \longrightarrow Cl_2(g) + 2H_2O(l)$	+1.64

(a) Give the conventional representation of the cell that is used to measure the standard electrode potential of iron as shown in the table.

.....

(2)

(b) With reference to electrons, give the meaning of the term **reducing agent**.

.....

.....

(1)

(c) Identify the weakest reducing agent from the species in the table.

Explain how you deduced your answer.

Species.....

Explanation.....

.....

(2)

(d) When HOCl acts as an oxidising agent, one of the atoms in the molecule is reduced.

(i) Place a tick (✓) next to the atom that is reduced.

Atom that is reduced	Tick (✓)
H	
O	
Cl	

(1)

(ii) Explain your answer to part (i) in terms of the change in the oxidation state of this atom.

.....

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(1)

(e) Using the information given in the table, deduce an equation for the redox reaction that would occur when hydroxide ions are added to HOCl

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(2)

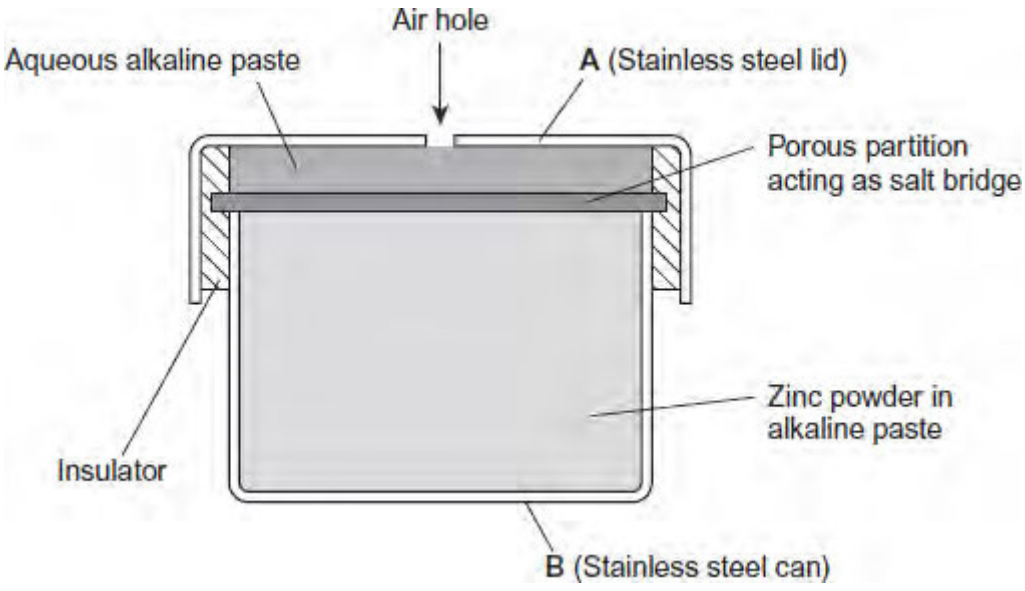
(f) The table is repeated to help you answer this question.

	$E^{\ominus}$ / V
$ZnO(s) + H_2O(l) + 2e^- \longrightarrow Zn(s) + 2OH^-(aq)$	–1.25

$Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$	–0.44
$O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$	+0.40
$2HOCl(aq) + 2H^+(aq) + 2e^- \longrightarrow Cl_2(g) + 2H_2O(l)$	+1.64

The half-equations from the table that involve zinc and oxygen are simplified versions of those that occur in hearing aid cells.

A simplified diagram of a hearing aid cell is shown in the following figure.



(i) Use data from the table to calculate the e.m.f. of this cell.

.....

.....

Answer = ...

(1)

(ii) Use half-equations from the table to construct an overall equation for the cell reaction.

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(1)

(iii) Identify which of **A** or **B**, in the figure, is the positive electrode. Give a reason for your answer.

Positive electrode .....

Reason .....

.....

.....

(2)

(iv) Suggest **one** reason, other than cost, why this type of cell is **not** recharged.

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(1)

(Total 14 marks)

Q5. Fuel cells are an increasingly important energy source for vehicles. Standard electrode potentials are used in understanding some familiar chemical reactions including those in fuel cells.

The following table contains some standard electrode potential data.

Electrode half-equation	$E^{\ominus}$ / V
$F_2 + 2e^- \longrightarrow 2F^-$	+2.87
$Cl_2 + 2e^- \longrightarrow 2Cl^-$	+1.36
$O_2 + 4H^+ + 4e^- \longrightarrow 2H_2O$	+1.23
$Br_2 + 2e^- \longrightarrow 2Br^-$	+1.07
$I_2 + 2e^- \longrightarrow 2I^-$	+0.54
$O_2 + 2H_2O + 4e^- \longrightarrow 4OH^-$	+0.40
$SO_4^{2-} + 4H^+ + 2e^- \longrightarrow SO_2 + 2H_2O$	+0.17
$2H^+ + 2e^- \longrightarrow H_2$	0.00
$4H_2O + 4e^- \longrightarrow 4OH^- + 2H_2$	−0.83

(a) A salt bridge was used in a cell to measure electrode potential.

Explain the function of the salt bridge.

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(2)

(b) Use data from the table above to deduce the halide ion that is the weakest reducing agent.

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(1)

(c) Use data from the table to justify why sulfate ions should **not** be capable of oxidising bromide ions.

.....

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.....

(1)

(d) Use data from the table to calculate a value for the EMF of a hydrogen–oxygen fuel cell operating under alkaline conditions.

EMF = ... V

(1)

(e) There are two ways to use hydrogen as a fuel for cars. One way is in a fuel cell to power an electric motor, the other is as a fuel in an internal combustion engine.

Suggest the major advantage of using the fuel cell.

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(1)
(Total 6 marks)

Q6. The following cell has an EMF of +0.46 V.

$$Cu \mid Cu^{2+} \parallel Ag^{+} \mid Ag$$

Which statement is correct about the operation of the cell?

A Metallic copper is oxidised by Ag^+ ions. ⬭

B The silver electrode has a negative polarity. ⬭

C The silver electrode gradually dissolves to form Ag^+ ions. ⬭

D Electrons flow from the silver electrode to the copper electrode via an external circuit. ⬭

(Total 1 mark)

Q1. 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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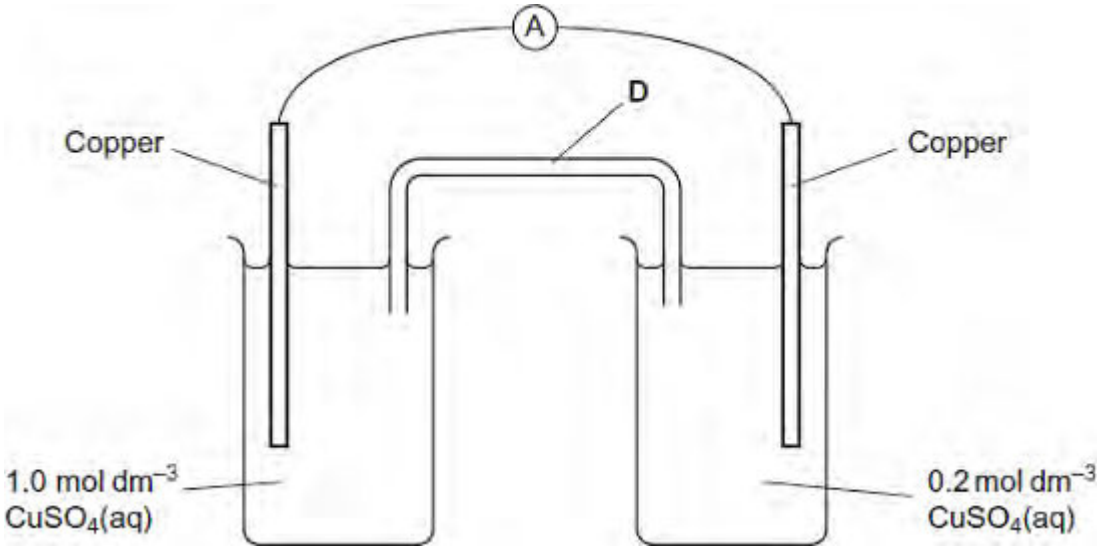
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(3)
(Total 8 marks)

Q2. 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.

.....

.....

(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)

Q3. 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*.

.....

.....

(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.

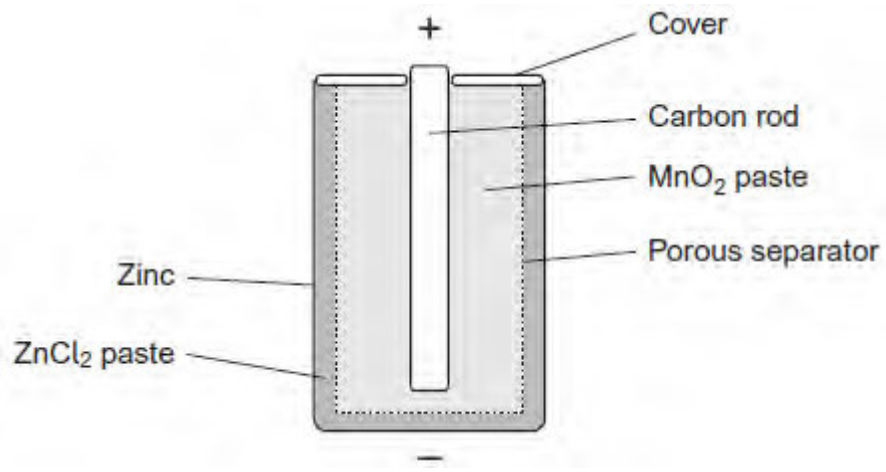
Weakest oxidising agent .....

Explanation .....

.....

(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.

.....

.....

(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).

.....

.....

(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)

Q4. 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)

Q5. 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 .....

.....

.....

(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)

Q1.(a) Use data from the table below to explain why dilute hydrochloric acid cannot be used to acidify potassium manganate(VII) in a titration.

	$E^{\ominus}$ / V
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$	+1.51
$Cl_2(aq) + 2e^- \rightarrow 2Cl^-(aq)$	+1.36
$2H^+(aq) + 2e^- \rightarrow H_2(aq)$	0.00

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(2)

(b) Use information from the table in part (a) to determine the minimum volume, in cm^3, of 0.500 mol dm^{-3} sulfuric acid that is required for a titre of 25.0 cm^3 of 0.0200 mol dm^{-3} potassium manganate(VII) solution.
Show your working.

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(3)

(c) In each titration using potassium manganate(VII), a large excess of dilute sulfuric acid is used to avoid any possibility of the brown solid MnO_2 forming.

(i) Deduce a half-equation for the reduction of MnO_4^- ions in acidic solution to form MnO_2.

.....

.....

(1)

(ii) Give **two** reasons why it is essential to avoid this reaction in a titration between potassium manganate(VII) and iron(II) ions.

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(2)

(d) Potassium manganate(VII) is an oxidising agent.
Suggest **one** reason why a 0.0200 mol dm^{-3} solution of potassium manganate(VII) does **not** need to be kept away from flammable material.

.....

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(1)
(Total 9 marks)

Q2. The table below shows some standard electrode potentials.

			$E^{\ominus}$ / V
$MnO_4^- + 8H^+ + 5e^-$	⟶	$Mn^{2+} + 4H_2O$	+1.51
$Cl_2(g) + 2e^-$	⟶	$2Cl^-(aq)$	+1.36
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	⟶	$2Cr^{3+} + 7H_2O$	+1.33

A student determined the concentration of iron(II) ions in a solution of iron(II) chloride by titration with acidified potassium dichromate(VI) solution. A second student titrated the same solution of iron(II) chloride with acidified potassium manganate(VII) solution.
By reference to the table, explain why the second student obtained a greater value for the concentration of iron(II) ions.

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(Total 2 marks)

Q3. (a) Lithium ion cells are used to power cameras and mobile phones.
A simplified representation of a cell is shown below.

$$Li \mid Li^+ \parallel Li^+, CoO_2 \mid LiCoO_2 \mid Pt$$

The reagents in the cell are absorbed onto powdered graphite that acts as a support medium. The support medium allows the ions to react in the absence of a solvent such as water.

The half-equation for the reaction at the positive electrode can be represented as follows.

$$Li^+ + CoO_2 + e^- \longrightarrow Li^+[CoO_2]^-$$

(i) Identify the element that undergoes a change in oxidation state at the positive electrode and deduce these oxidation states of the element.

Element .....

Oxidation state 1 .....

Oxidation state 2 .....

.....

(3)

(ii) Write a half-equation for the reaction at the negative electrode during operation of the lithium ion cell.

.....

(1)

(iii) Suggest two properties of platinum that make it suitable for use as an external electrical contact in the cell.

Property 1 .....

Property 2 .....

(2)

(iv) Suggest **one** reason why water is **not** used as a solvent in this cell.

.....

.....

(1)

(b) The half-equations for two electrodes used to make an electrochemical cell are shown below.

$ClO_3^-(aq) + 6H^+(aq) + 6e^- \longrightarrow Cl^-(aq) + 3H_2O(l)$ $\quad E^\ominus = +1.45$ V

$SO_4^{2-}(aq) + 2H^+(aq) + 2e^- \longrightarrow SO_3^{2-}(aq) + H_2O(l)$ $\quad E^\ominus = +0.17$ V

(i) Write the conventional representation for the cell using platinum contacts.

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(2)

(ii) Write an overall equation for the cell reaction and identify the oxidising and reducing agents.

Overall equation .....

.....

.....

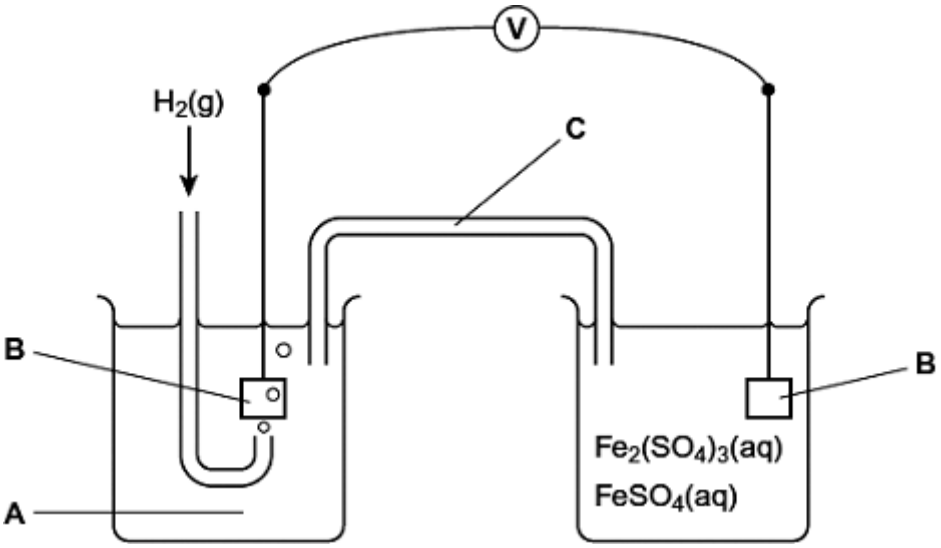
Oxidising agent .....

Reducing agent .....

(3)

(Total 12 marks)

Q4. The diagram below shows a cell that can be used to measure the standard electrode potential for the half-reaction $Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq)$. In this cell, the electrode on the right-hand side is positive.



(a) Identify solution **A** and give its concentration. State the other essential conditions for the operation of the standard electrode that forms the left-hand side of the cell.

Solution **A** .....

Conditions .....

.....

.....

(3)

(b) Identify the material from which electrodes **B** are made. Give **two** reasons why this material is suitable for its purpose.

Material .....

Reason 1 .....

.....

Reason 2 .....

.....

(3)

(c) Identify a solution that could be used in **C** to complete the circuit. Give **two** reasons why this solution is suitable for its purpose.

Solution .....

Reason 1 .....

.....

Reason 2 .....

.....

(3)

(d) Write the conventional representation for this cell.

.....

(1)

(e) The voltmeter **V** shown in the diagram of the cell was replaced by an ammeter.

(i) Write an equation for the overall cell reaction that would occur.

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(1)

(ii) Explain why the ammeter reading would fall to zero after a time.

.....

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(1)

(Total 12 marks)

Q5. 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.

.....

(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.

.....

.....

(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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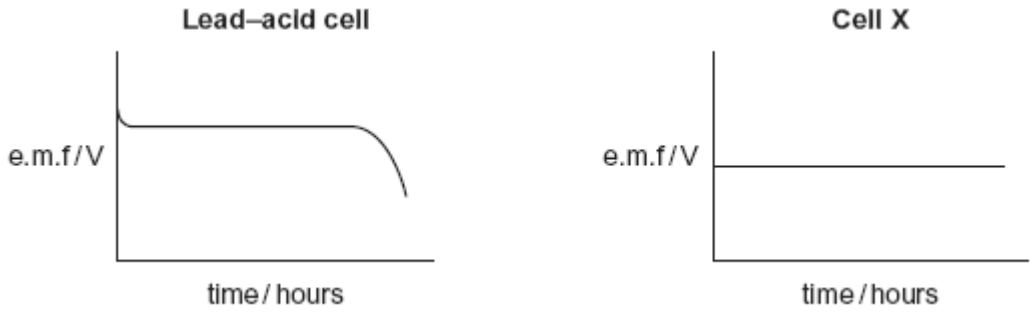
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(2)

(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)

Q6. Some electrode potentials are shown in the table below. These values are **not** listed in numerical order.

Electrode half-equation	$E^{\ominus}$ / V
$Cl_2(aq) + 2e^- \longrightarrow 2Cl^-(aq)$	+1.36
$2HOCl(aq) + 2H^+(aq) + 2e^- \longrightarrow Cl_2(aq) + 2H_2O(l)$	+1.64
$H_2O_2(aq) + 2H^+(aq) + 2e^- \longrightarrow 2H_2O(l)$	+1.77
$O_2(g) + 2H^+(aq) + 2e^- \longrightarrow H_2O_2(aq)$	+0.68
$O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l)$	+1.23

(a) Identify the most powerful reducing agent from all the species in the table.

.....

(1)

(b) Use data from the table to explain why chlorine should undergo a redox reaction with water. Write an equation for this reaction.

Explanation .....

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.....

Equation .....

.....

.....

(2)

(c) Suggest **one** reason why the redox reaction between chlorine and water does not normally occur in the absence of light.

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(1)

(d) Use the appropriate half-equation from the table to explain in terms of oxidation states what happens to hydrogen peroxide when it is reduced.

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(2)

(e) Use data from the table to explain why one molecule of hydrogen peroxide can oxidise another molecule of hydrogen peroxide. Write an equation for the redox reaction that occurs.

Explanation .....

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Equation .....

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(2)
(Total 8 marks)

Q1. The electrons transferred in redox reactions can be used by electrochemical cells to provide energy.

Some electrode half-equations and their standard electrode potentials are shown in the table below.

Half-equation	$E^{\ominus}$/V
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$	+1.33
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.77
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$	–0.44
$Li^+(aq) + e^- \rightarrow Li(s)$	–3.04

(a) Describe a standard hydrogen electrode.

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(4)

(b) A conventional representation of a lithium cell is given below.
This cell has an e.m.f. of +2.91 V

$$Li(s) \mid Li^+(aq) \parallel Li^+(aq) \mid MnO_2(s) , LiMnO_2(s) \mid Pt(s)$$

Write a half-equation for the reaction that occurs at the positive electrode of this cell.

Calculate the standard electrode potential of this positive electrode.

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(2)

(c) Suggest what reactions occur, if any, when hydrogen gas is bubbled into a solution containing a mixture of iron(II) and iron(III) ions. Explain your answer.

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(2)

(d) A solution of iron(II) sulfate was prepared by dissolving 10.00 g of $FeSO_4.7H_2O$ (M_r = 277.9) in water and making up to 250 cm^3 of solution. The solution was left to stand, exposed to air, and some of the iron(II) ions became oxidised to iron(III) ions. A 25.0 cm^3 sample of the partially oxidised solution required 23.70 cm^3 of 0.0100 mol dm^{-3} potassium dichromate(VI) solution for complete reaction in the presence of an excess of dilute sulfuric acid.

Calculate the percentage of iron(II) ions that had been oxidised by the air.

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(6)
(Total 14 marks)

Q2. In this question consider the data below.

	$E^{\ominus}$ / V
$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0.80
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$	−0.13

The e.m.f. of the cell $Ag(s) \mid Ag^+(aq) \parallel Pb^{2+}(aq) \mid Pb(s)$ is

A 0.93 V

B 0.67 V

C −0.67 V

D −0.93 V

(Total 1 mark)

Q3. Nickel–cadmium cells are used to power electrical equipment such as drills and shavers.
The electrode reactions are shown below.

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

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

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

.....

(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)

Q4. (a) The term oxidation was used originally to describe a reaction in which a substance gained oxygen. The oxygen was provided by the oxidising agent. Later the definition of oxidation was revised when the importance of electron transfer was recognised.

An aqueous solution of sulfur dioxide was reacted in separate experiments as follows.

Reaction **1** with HgO

$$H_2O + SO_2 + HgO \rightarrow H_2SO_4 + Hg$$

Reaction **2** with chlorine

$$2H_2O + SO_2 + Cl_2 \rightarrow H_2SO_4 + 2HCl$$

(i) In Reaction **1**, identify the substance that donates oxygen and therefore is the oxidising agent.

.....

(ii) Show, by writing a half-equation, that this oxidising agent in reaction **1** is an electron acceptor.

.....

(iii) Write a half-equation for the oxidation process occurring in reaction **2**.

.....

(iv) Write a half-equation for the reduction process occurring in reaction **2**.

.....

(4)

(b) Use the standard electrode potential data given in the able below to answer the questions which follow.

	E / V
$V^{3+}(aq) + e^- \rightarrow V^{2+}(aq)$	–0.26
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow H_2SO_3(aq) + H_2O(l)$	+0.17
$VO^{2+}(aq) + 2H^+(aq) + e^- \rightarrow V^{3+}(aq) + H_2O(l)$	+0.34
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.77
$VO_2^+(aq) + 2H^+(aq) + e^- \rightarrow VO^{2+}(aq) + H_2O(l)$	+1.00
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$	+1.52

Each of the above can be reversed under suitable conditions

(i) An excess of potassium manganate(VII) was added to a solution containing $V^{2+}(aq)$ ions. Determine the vanadium species present in the solution at the end of this reaction. State the oxidation state of vanadium in this species and write a half-equation for its formation from $V^{2+}(aq)$.

Vanadium species present at the end of the reaction .....

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Oxidation state of vanadium in the final species .....

.....

Half-equation .....

(ii) The cell represented below was set up under standard conditions.

$$Pt|H_2SO_3(aq),SO_4^{2-}(aq),H^+(aq)||Fe^{3+}(aq),Fe^{2+}(aq)|Pt$$

Calculate the e.m.f. of this cell and state, with an explanation, how this e.m.f. will change if the concentration of $Fe^{3+}(aq)$ ions is increased.

Cell e.m.f. .....

Change in cell e.m.f. .....

Explanation .....

.....

(7)

(c) Consider the cell below

$$\overset{-}{Pt|H_2(g)|H^+(aq)||O_2(g)|OH^-(aq)|Pt}\overset{+}{}$$

(i) Using half-equations, deduce an overall equation for the cell reaction.

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.....

(ii) State how, if at all, the e.m.f. of this cell will change if the surface area of each platinum electrode is doubled.

.....

(3)

(d) Currently, almost all hydrogen is produced by the high-temperature reaction between methane, from North Sea gas, and steam. Give one economic and one environmental disadvantage of this method of producing hydrogen.

Economic disadvantage .....

Environmental disadvantage .....

(2)

(e) Hydrogen can also be produced by the electrolysis of acidified water using electricity produced using solar cells. Give one reason why this method is not used on a large scale.

.....

(1)

(Total 17 marks)

Q5. Use the standard electrode potential data given in the table below, where appropriate, to answer the questions which follow.

	$E^{\ominus}$/V
$V^{3+}(aq) + e^- \rightarrow V^{2+}(aq)$	–0.26
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow H_2SO_3(aq) + H_2O$	+0.17
$VO^{2+}(aq) + 2H^+(aq) + e^- \rightarrow V^{3+}(aq) + H_2O(l)$	+0.34
$O_2(g) + 2H^+(aq) + 2e^- \rightarrow H_2O_2(aq)$	+0.68
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.77
$VO_2^+(aq) + 2H^+(aq) + e^- \rightarrow VO^{2+}(aq) + H_2O(l)$	+1.00
$2IO_3^-(aq) + 12H^+(aq) + 10e^- \rightarrow I_2(aq) + 6H_2O(l)$	+1.19
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$	+1.52

Each of the above can be reversed under suitable conditions.

(a) The cell represented below was set up under standard conditions.

$Pt \mid H_2SO_3(aq), SO_4^{2-}(aq), \parallel Fe^{3+}(aq), Fe^{2+}(aq) \mid Pt$

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

.....

(ii) Write a half-equation for the oxidation process occurring at the negative electrode of this cell.

.....

(2)

(b) The cell represented below was set up under standard conditions.

$Pt \mid H_2O_2(aq), O_2(g) \parallel IO_3^-(aq), I_2(aq) \mid Pt$

(i) Write an equation for the spontaneous cell reaction.

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.....

(ii) Give **one** reason why the e.m.f. of this cell changes when the electrodes are connected and a current flows.

.....

(iii) State how, if at all, the e.m.f. of this standard cell will change if the surface area of each platinum electrode is doubled.

.....

(iv) State how, if at all, the e.m.f. of this cell will change if the concentration of IO_3^- ions is increased. Explain your answer.

Change, if any, in e.m.f. of cell .....

Explanation .....

.....

(7)

(c) An excess of acidified potassium manganate(VII) was added to a solution containing $V^{2+}(aq)$ ions. Use the data given in the table to determine the vanadium species present in the solution at the end of this reaction. State the oxidation state of vanadium in this species and write a half-equation for its formation from $V^{2+}(aq)$.

Vanadium species present at end of reaction .....

Oxidation state of vanadium in final species ….....

Half-equation .....

(3)

(Total 12 marks)

Q6. Use the data in the table below, where appropriate, to answer the questions which follow.

Standard electrode potentials	$E^{\ominus}$ / V
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.77
$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$	+1.36
$2BrO_3^-(aq) + 12H^+(aq) + 10e^- \rightarrow Br_2(aq) + 6H_2O(l)$	+1.52
$O_3(g) + 2H^+(aq) + 2e^- \rightarrow O_2(g) + H_2O(l)$	+2.08
$F_2O(g) + 2H^+(aq) + 4e^- \rightarrow 2F^-(aq) + H_2O(l)$	+2.15

Each of the above can be reversed under suitable conditions.

(a) (i) Identify the most powerful reducing agent in the table.

.....

(ii) Identify the most powerful oxidising agent in the table.

.....

(iii) Identify **all** the species in the table which can be oxidised in acidic solution by $BrO_3^-(aq)$.

.....

(4)

(b) The cell represented below was set up.

$Pt|Fe^{2+}(aq), Fe^{3+}(aq) || BrO_3^-(aq), Br_2(aq)|Pt$

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

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(ii) Write a half-equation for the reaction occurring at the negative electrode when current is taken from this cell.

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(iii) Deduce what change in the concentration of $Fe^{3+}(aq)$ would cause an increase in the e.m.f. of the cell. Explain your answer.

Change in concentration .....

Explanation .....

.....

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(6)
(Total 10 marks)

Q1. Where appropriate, use the standard electrode potential data in the table below to answer the questions which follow.

							$E^{\ominus}$/V
$Zn^{2+}(aq)$	+	$2e^-$	→	$Zn(s)$			–0.76
$V^{3+}(aq)$	+	e^-	→	$V^{2+}(aq)$			–0.26
$SO_4^{2-}(aq) + 2H^+(aq)$	+	$2e^-$	→	$SO_3^{2-}(aq)$	+	$H_2O(l)$	+0.17
$VO^{2+}(aq) + 2H^+(aq)$	+	e^-	→	$V^{3+}(aq)$	+	$H_2O(l)$	+0.34
$Fe^{3+}(aq)$	+	e^-	→	$Fe^{2+}(aq)$			+0.77
$VO_2^+(aq) + 2H^+(aq)$	+	e^-	→	$VO^{2+}(aq)$	+	$H_2O(l)$	+1.00
$Cl_2(aq)$	+	$2e^-$	→	$2Cl^-(aq)$			+1.36

(a) From the table above select the species which is the most powerful reducing agent.

.....

(1)

(b) From the table above select

(i) a species which, in acidic solution, will reduce $VO_2^+(aq)$ to $VO^{2+}(aq)$ but will **not** reduce $VO^{2+}(aq)$ to $V^{3+}(aq)$,

.....

(ii) a species which, in acidic solution, will oxidise $VO^{2+}(aq)$ to $VO_2^+(aq)$.

.....

(2)

(c) The cell represented below was set up under standard conditions.

$Pt|Fe^{2+}(aq), Fe^{3+}(aq)||Tl^{3+}(aq), Tl^+(aq)|Pt$ Cell e.m.f. = + 0.48 V

(i) Deduce the standard electrode potential for the following half-reaction.

$Tl^{3+}(aq) + 2e^- \rightarrow Tl^+(aq)$

.....

.....

(ii) Write an equation for the spontaneous cell reaction.

.....

(3)

(d) After acidification, 25.0 cm^3 of a solution of hydrogen peroxide reacted exactly with 16.2 cm^3 of a 0.0200 mol dm^{-3} solution of potassium manganate(VII). The overall equation for the reaction is given below.

$2MnO_4^- + 6H^+ + 5H_2O_2 \rightarrow 2Mn^{2+} + 8H_2O + 5O_2$

(i) Use the equation for this reaction to determine the concentration, in g dm^{-3}, of the hydrogen peroxide solution.

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(ii) Calculate the maximum volume of oxygen, measured at a pressure of 98 kPa and a temperature of 298 K, which would be evolved in this reaction.

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(8)
(Total 14 marks)

Q2. Large blocks of magnesium are bolted onto the hulls of iron ships in an attempt to prevent the iron being converted into iron(II), one of the steps in the rusting process.

Use the data below, where appropriate, to answer the questions which follow.

			$E^{\ominus}$ / V
$Mg^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Mg(s)$	–2.37
$Fe^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Fe(s)$	–0.44
$O_2(g) + 2H_2O(l) + 4e^-$	$\rightleftharpoons$	$4OH^-(aq)$	+0.40

(a) Calculate the e.m.f. of the cell represented by $Mg(s)|Mg^{2+}(aq)||Fe^{2+}(aq)|Fe(s)$ under standard conditions. Write a half-equation for the reaction occurring at the negative electrode of this cell when a current is drawn.

Cell e.m.f. .....

Half-equation .....

.....

(2)

(b) Deduce how the e.m.f. of the cell $Mg(s)|Mg^{2+}(aq)||Fe^{2+}(aq)|Fe(s)$ changes when the concentration of Mg^{2+} is decreased. Explain your answer.

Change in e.m.f. .....

Explanation .....

.....

(3)

(c) Calculate a value for the e.m.f. of the cell represented by $Pt(s)|OH^-(aq)|O_2(g)||Fe^{2+}(aq)|Fe(s)$ and use it to explain why iron corrodes when in contact with water which contains dissolved oxygen.

Cell e.m.f. .....

Explanation .....

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(2)
(Total 7 marks)

Q3. Use the standard electrode potential data in the table below to answer the questions which follow.

			$E^{\ominus}$ / V
$Ce^{4+}(aq) + e^-$	⇌	$Ce^{3+}(aq)$	+1.70
$MnO^-(aq) + 8H^+(aq) + 5e^-$	⇌	$Mn^{2+}(aq) + 4H_2O(l)$	+1.51
$Cl_2(g) + 2e^-$	⇌	$2Cl^-(aq)$	+1.36
$VO_2^+(aq) + 2H^+(aq) + e^-$	⇌	$VO^{2+}(aq) + H_2O(l)$	+1.00
$Fe^{3+}(aq) + e^-$	⇌	$Fe^{2+}(aq)$	+0.77
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$	⇌	$H_2SO_3(aq) + H_2O(l)$	+0.17

(a) Name the standard reference electrode against which all other electrode potentials are measured.

.....

(1)

(b) When the standard electrode potential for $Fe^{3+}(aq) / Fe^{2+}(aq)$ is measured, a platinum electrode is required.

(i) What is the function of the platinum electrode?

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(ii) What are the standard conditions which apply to $Fe^{3+}(aq)/Fe^{2+}(aq)$ when measuring this potential?

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(3)

(c) The cell represented below was set up under standard conditions.

$$Pt|H_2SO_3(aq), SO_4^{2-}(aq)||MnO_4^-(aq), Mn^{2+}(aq)|Pt$$

Calculate the e.m.f. of this cell and write an equation for the spontaneous cell reaction.

Cell e.m.f. .....

Equation .....

.....

(3)

(d) (i) Which one of the species given in the table is the strongest oxidising agent?

.....

(ii) Which of the species in the table could convert $Fe^{2+}(aq)$ into $Fe^{3+}(aq)$ but could not convert $Mn^{2+}(aq)$ into $MnO_4^-(aq)$?

.....

(3)

(e) Use data from the table of standard electrode potentials to deduce the cell which would have a standard e.m.f. of 0.93 V. Represent this cell using the convention shown in part (c).

.....

(2)
(Total 12 marks)

Q4. Use the data in the table below to answer this question.

	$E^{\ominus}$ / V
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$	+ 1.52
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$	+ 1.33
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+ 0.77

$Cr^{3+}(aq) + e^- \rightarrow Cr^{2+}(aq)$	− 0.41
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	− 0.76

Which one of the following statements is **not** correct?

A $Fe^{2+}(aq)$ can reduce acidified $MnO_4^-(aq)$ to $Mn^{2+}(aq)$

B $CrO_7^{2-}(aq)$ can oxidise acidified $Fe^{2+}(aq)$ to $Fe^{3+}(aq)$

C $Zn(s)$ can reduce acidified $Cr_2O_7^{2-}(aq)$ to $Cr^{2+}(aq)$

D $Fe^{2+}(aq)$ can reduce acidified $Cr^{3+}(aq)$ to $Cr^{2+}(aq)$

(Total 1 mark)

Q5. Use the data in the table below to answer this question.

	$E^{\ominus}$ / V
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$	+ 1.52
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$	+ 1.33
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+ 0.77
$Cr^{3+}(aq) + e^- \rightarrow Cr^{2+}(aq)$	− 0.41
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	− 0.76

The most powerful oxidising agent in the table is

A $Mn^{2+}(aq)$

B $Zn(s)$

C $MnO_4^-(aq)$

D $Zn^{2+}(aq)$

(Total 1 mark)

Q6. In this question consider the data below.

	$E^{\ominus}$ / V
$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0.80
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$	−0.13

The e.m.f. of the cell $Pt(s) \mid H_2(g) \mid H^+(aq) \parallel Ag^+(aq) \mid Ag(s)$ would be increased by

A increasing the concentration of $H^+(aq)$.

B increasing the surface area of the Pt electrode.

C increasing the concentration of $Ag^+(aq)$.

D decreasing the pressure of $H_2(g)$.

(Total 1 mark)

Q7. Use the data below, where appropriate, to answer the questions which follow.

Standard electrode potentials	$E^{\ominus}$/V
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$Br_2(aq) + 2e^- \rightarrow 2Br^-(aq)$	+1.09
$2BrO_3^-(aq) + 12H^+(aq) + 10e^- \rightarrow Br_2(aq) + 6H_2O(l)$	+1.52

Each of the above can be reversed under suitable conditions.

(a) State the hydrogen ion concentration and the hydrogen gas pressure when, at 298 K, the potential of the hydrogen electrode is 0.00 V.

Hydrogen ion concentration .....

Hydrogen gas pressure .....

(2)

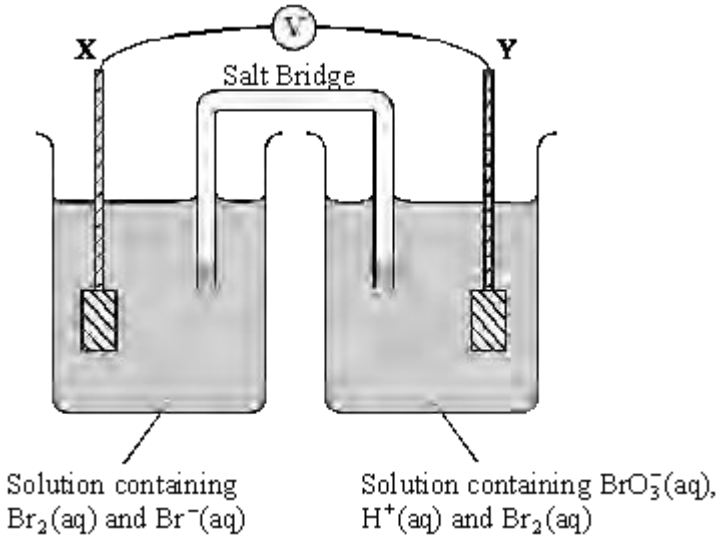
(b) The electrode potential of a hydrogen electrode changes when the hydrogen ion concentration is reduced. Explain, using Le Chatelier's principle, why this change occurs and state how the electrode potential of the hydrogen electrode changes.

Explanation of change .....

Change in electrode potential .....

(3)

(c) A diagram of a cell using platinum electrodes **X** and **Y** is shown below.



(i) Use the data above to calculate the e.m.f. of the above cell under standard conditions.

.....

(ii) Write a half-equation for the reaction occurring at electrode **X** and an overall equation for the cell reaction which occurs when electrodes **X** and **Y** are connected.

Half-equation .....

Overall equation .....

.....

(4)
(Total 9 marks)

Q8.

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l) \quad E^{\ominus} = +1.33\ V$$

$$Br_2(aq) + 2e^- \rightarrow 2Br^-(aq) \quad E^{\ominus} = +1.09\ V$$

$$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq) \quad E^{\ominus} = +0.77\ V$$

$$VO^{2+}(aq) + 2H^+(aq) + e^- \rightarrow V^{3+}(aq) + H_2O(l) \quad E^{\ominus} = +0.34\ V$$

$$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow H_2SO_3(aq) + H_2O(l) \quad E^{\ominus} = +0.17\ V$$

Based on the above data, which one of the following could reduce 0.012 mol of bromine to bromide ions?

A 40 cm^3 of a 0.10 mol dm^{-3} solution of $Cr_2O_7^{2-}(aq)$

B 80 cm^3 of a 0.30 mol dm^{-3} solution of $Fe^{3+}(aq)$

C 50 cm^3 of a 0.24 mol dm^{-3} solution of $V^{3+}(aq)$

C 50 cm^3 of a 0.24 mol dm^{-3} solution of $H_2SO_3(aq)$

(Total 1 mark)

Q9. A disproportionation reaction occurs when a species $\mathbf{M}^+$ spontaneously undergoes simultaneous oxidation and reduction.

$$2M^+(aq) \rightarrow M^{2+}(aq) + M(s)$$

The table below contains $E^{\ominus}$ data for copper and mercury species.

	$E^{\ominus}$/ V
$Cu^{2+}(aq) + e^- \rightarrow Cu^+(aq)$	+ 0.15
$Cu^+(aq) + e^- \rightarrow Cu(s)$	+ 0.52
$Hg^{2+}(aq) + e^- \rightarrow Hg^+(aq)$	+ 0.91
$Hg^+(aq) + e^- \rightarrow Hg(l)$	+ 0.80

Using these data, which one of the following can be predicted?

A Both Cu(I) and Hg(I) undergo disproportionation.

B Only Cu(I) undergoes disproportionation.

C Only Hg(I) undergoes disproportionation.

D Neither Cu(I) nor Hg(I) undergoes disproportionation.

(Total 1 mark)

Q1. The following cell has an EMF of +0.46 V.

$$Cu \mid Cu^{2+} \parallel Ag^{+} \mid Ag$$

Which statement is correct about the operation of the cell?

A Metallic copper is oxidised by Ag^+ ions.

B The silver electrode has a negative polarity.

C The silver electrode gradually dissolves to form Ag^+ ions.

D Electrons flow from the silver electrode to the copper electrode via an external circuit.

(Total 1 mark)

Q2. In this question consider the data below.

	$E^{\ominus}$ / V
$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0.80
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$	−0.13

The e.m.f. of the cell $Ag(s) \mid Ag^+(aq) \parallel Pb^{2+}(aq) \mid Pb(s)$ is

A 0.93 V

B 0.67 V

C −0.67 V

D −0.93 V

(Total 1 mark)

Q3. In this question consider the data below.

	$E^{\ominus}$ / V
$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0.80
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$	−0.13

The e.m.f. of the cell $Pt(s) \mid H_2(g) \mid H^+(aq) \mid\mid Ag^+(aq) \mid Ag(s)$ would be increased by

A increasing the concentration of $H^+(aq)$.

B increasing the surface area of the Pt electrode.

C increasing the concentration of $Ag^+(aq)$.

D decreasing the pressure of $H_2(g)$.

(Total 1 mark)

Q4. A disproportionation reaction occurs when a species M^+ spontaneously undergoes simultaneous oxidation and reduction.

$$2M^+(aq) \rightarrow M^{2+}(aq) + M(s)$$

The table below contains $E^{\ominus}$ data for copper and mercury species.

	$E^{\ominus}$ / V
$Cu^{2+}(aq) + e^- \rightarrow Cu^+(aq)$	+ 0.15
$Cu^+(aq) + e^- \rightarrow Cu(s)$	+ 0.52
$Hg^{2+}(aq) + e^- \rightarrow Hg^+(aq)$	+ 0.91
$Hg^+(aq) + e^- \rightarrow Hg(l)$	+ 0.80

Using these data, which one of the following can be predicted?

A Both Cu(I) and Hg(I) undergo disproportionation.

B Only Cu(I) undergoes disproportionation.

C Only Hg(I) undergoes disproportionation.

D Neither Cu(I) nor Hg(I) undergoes disproportionation.

(Total 1 mark)

Q5.

$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$ $\quad E^{\ominus} = +1.33$ V

$Br_2(aq) + 2e^- \rightarrow 2Br^-(aq)$ $\quad E^{\ominus} = +1.09$ V

$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$ $\quad E^{\ominus} = +0.77$ V

$VO^{2+}(aq) + 2H^+(aq) + e^- \rightarrow V^{3+}(aq) + H_2O(l)$ $\quad E^{\ominus} = +0.34$ V

$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow H_2SO_3(aq) + H_2O(l)$ $\quad E^{\ominus} = +0.17$ V

Based on the above data, which one of the following could reduce 0.012 mol of bromine to bromide ions?

A 40 cm^3 of a 0.10 mol dm^{-3} solution of $Cr_2O_7^{2-}(aq)$

B 80 cm^3 of a 0.30 mol dm^{-3} solution of $Fe^{3+}(aq)$

C 50 cm^3 of a 0.24 mol dm^{-3} solution of $V^{3+}(aq)$

C 50 cm^3 of a 0.24 mol dm^{-3} solution of $H_2SO_3(aq)$

(Total 1 mark)

Q6. Use the data in the table below to answer this question.

	$E^{\ominus}$ / V
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$	+ 1.52
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$	+ 1.33
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+ 0.77
$Cr^{3+}(aq) + e^- \rightarrow Cr^{2+}(aq)$	− 0.41
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	− 0.76

The most powerful oxidising agent in the table is

A $Mn^{2+}(aq)$

B $Zn(s)$

C $MnO_4^-(aq)$

D $Zn^{2+}(aq)$

(Total 1 mark)

Q7. Use the data in the table below to answer this question.

	$E^{\ominus}$ / V
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$	+ 1.52
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$	+ 1.33
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+ 0.77
$Cr^{3+}(aq) + e^- \rightarrow Cr^{2+}(aq)$	− 0.41
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	− 0.76

Which one of the following statements is **not** correct?

A $Fe^{2+}(aq)$ can reduce acidified $MnO_4^-(aq)$ to $Mn^{2+}(aq)$

B $CrO_7^{2-}(aq)$ can oxidise acidified $Fe^{2+}(aq)$ to $Fe^{3+}(aq)$

C $Zn(s)$ can reduce acidified $Cr_2O_7^{2-}(aq)$ to $Cr^{2+}(aq)$

D $Fe^{2+}(aq)$ can reduce acidified $Cr^{3+}(aq)$ to $Cr^{2+}(aq)$

(Total 1 mark)