

AQA A-LEVEL CHEMISTRY • 7405

CHEMISTRY



EXAM QUESTION PACK

3.1.11 ELECTRODE POTENTIALS AND ELECTROCHEMICAL CELLS

SCORE

/463

PERCENTAGE

%

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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	$E^\ominus / \text{V}$
$\text{F}_2 + 2\text{e}^- \longrightarrow 2\text{F}^-$	+2.87
$\text{Cl}_2 + 2\text{e}^- \longrightarrow 2\text{Cl}^-$	+1.36
$\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \longrightarrow 2\text{H}_2\text{O}$	+1.23
$\text{Br}_2 + 2\text{e}^- \longrightarrow 2\text{Br}^-$	+1.07
$\text{I}_2 + 2\text{e}^- \longrightarrow 2\text{I}^-$	+0.54
$\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \longrightarrow 4\text{OH}^-$	+0.40
$\text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^- \longrightarrow \text{SO}_2 + 2\text{H}_2\text{O}$	+0.17
$2\text{H}^+ + 2\text{e}^- \longrightarrow \text{H}_2$	0.00
$4\text{H}_2\text{O} + 4\text{e}^- \longrightarrow 4\text{OH}^- + 2\text{H}_2$	-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.

(1)

(Total 6 marks)

Q2.

A representation of a hydrogen–oxygen fuel cell that operates in alkaline conditions is



- (a) (i) Write a half-equation for the reaction that occurs at each electrode. Use the half-equations to deduce an overall equation for the cell.

Half-equation at positive electrode _____

Half-equation at negative electrode _____

Overall equation _____

(3)

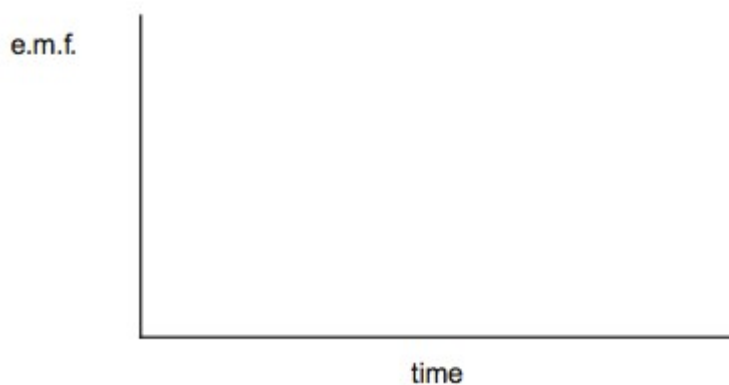
- (ii) State and explain the effect, if any, of increasing the pressure of oxygen on the e.m.f. of this cell.

Effect on e.m.f. _____

Explanation _____

(2)

- (b) Complete the diagram to show how the e.m.f. of a hydrogen–oxygen fuel cell changes with time.



(1)

- (c) (i) Suggest the effect, if any, on the e.m.f. of this cell if the surface area of each platinum electrode is increased.

(1)

- (ii) State the main environmental advantage of using a hydrogen–oxygen fuel cell to power a car.

(1)

- (d) Suggest why the use of a hydrogen–oxygen fuel cell might not be carbon-neutral.

(1)

(Total 9 marks)

Q3.

The table below shows some standard electrode potential data.

	$E^\ominus / \text{V}$
$\text{ZnO(s)} + \text{H}_2\text{O(l)} + 2\text{e}^- \longrightarrow \text{Zn(s)} + 2\text{OH}^-(\text{aq})$	–1.25
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Fe(s)}$	–0.44
$\text{O}_2(\text{g}) + 2\text{H}_2\text{O(l)} + 4\text{e}^- \longrightarrow 4\text{OH}^-(\text{aq})$	+0.40
$2\text{HOCl(aq)} + 2\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{Cl}_2(\text{g}) + 2\text{H}_2\text{O(l)}$	+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.

(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

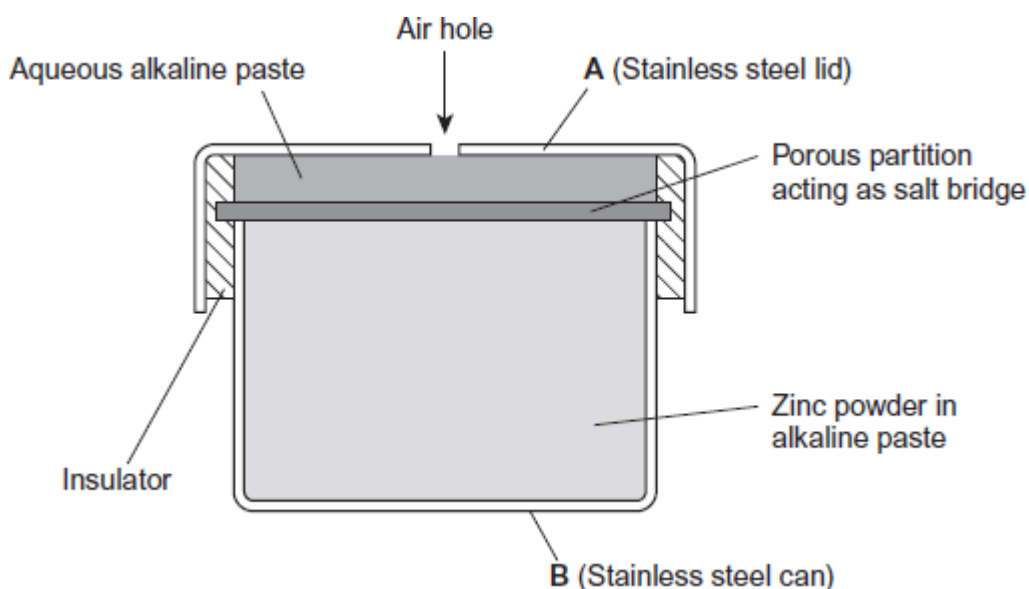
(2)

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

	$E^\ominus / \text{V}$
$\text{ZnO(s)} + \text{H}_2\text{O(l)} + 2\text{e}^- \longrightarrow \text{Zn(s)} + 2\text{OH}^-(\text{aq})$	-1.25
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Fe(s)}$	-0.44
$\text{O}_2(\text{g}) + 2\text{H}_2\text{O(l)} + 4\text{e}^- \longrightarrow 4\text{OH}^-(\text{aq})$	+0.40
$2\text{HOCl(aq)} + 2\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{Cl}_2(\text{g}) + 2\text{H}_2\text{O(l)}$	+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 _____

(2)

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

(1)

(Total 14 marks)

Q4.

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.

(1)

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

(1)

(Total 7 marks)

Q5.

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.

(2)

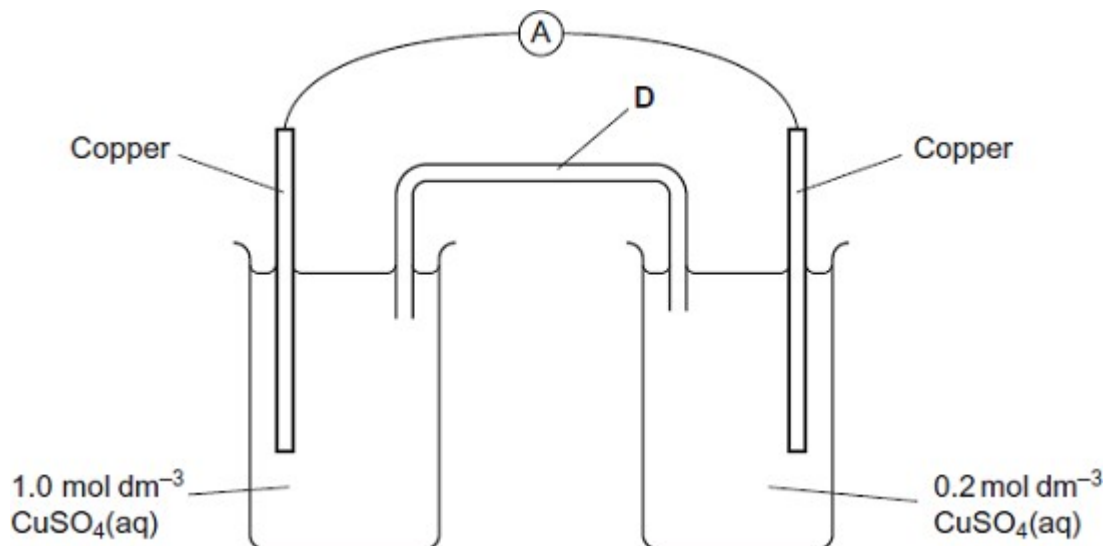
- (b) Suggest **one** reason why many waste disposal centres contain a separate section for cells and batteries.

(1)

(Total 3 marks)

Q6.

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.

(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** $\text{Li}^+ + \text{MnO}_2 + \text{e}^- \longrightarrow \text{LiMnO}_2$ $E = -0.15 \text{ V}$

Electrode **B** $\text{Li}^+ + \text{e}^- \longrightarrow \text{Li}$

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**.

(2)

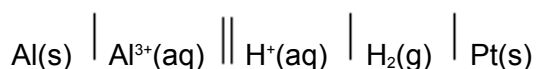
- (iii) Suggest why the recharging of a lithium cell may lead to release of carbon dioxide into the atmosphere.

(1)

(Total 9 marks)

Q7.

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



- (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.

(2)

(Total 5 marks)

Q8.

This table shows some standard electrode potential data.

Electrode half-equation	$E^\ominus / \text{V}$
$\text{Au}^+(\text{aq}) + \text{e}^- \longrightarrow \text{Au}(\text{s})$	+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 _____

(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.

(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.

(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.

(3)
(Total 15 marks)

Q9.

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*.

(1)

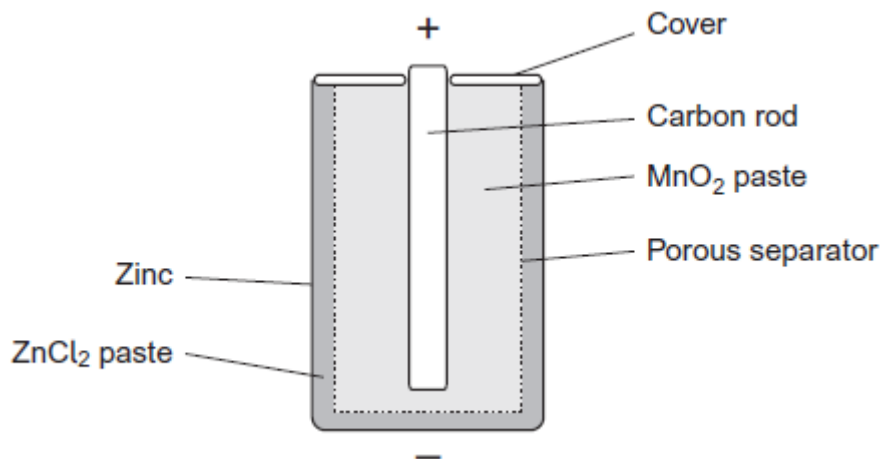
- (b) Deduce the identity of the weakest oxidising agent in the table.
Explain how $E^\ominus$ values can be used to make this deduction.

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.
- _____

(2)
(Total 17 marks)

Q10.

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	$E^\ominus / \text{V}$
$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn}(\text{s})$	-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.

(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	$E^\ominus / \text{V}$
$\text{PbO}_2(\text{s}) + 3\text{H}^+(\text{aq}) + \text{HSO}_4^-(\text{aq}) + 2\text{e}^- \longrightarrow \text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l})$	+1.69
$\text{PbSO}_4(\text{s}) + \text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{Pb}(\text{s}) + \text{HSO}_4^-(\text{aq})$	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.

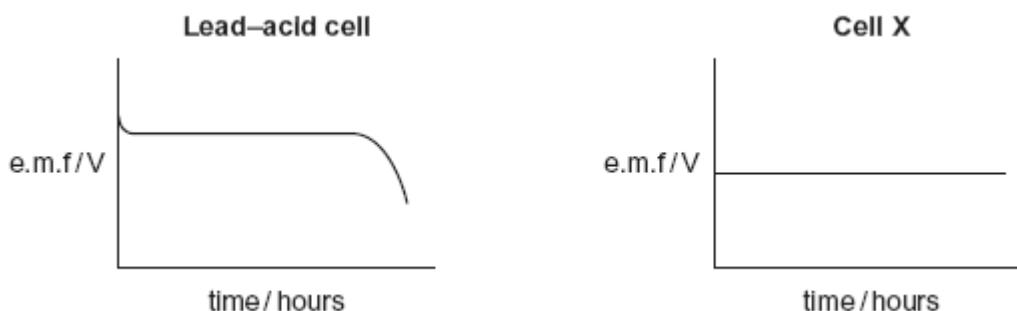
Use this information to calculate the missing electrode potential for the half-equation shown in the table.

(1)

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

(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**.

(2)

(Total 17 marks)

Q11.

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°/V
$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$	+1.33
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq})$	+0.77
$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	0.00
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Fe}(\text{s})$	-0.44
$\text{Li}^+(\text{aq}) + \text{e}^- \rightarrow \text{Li}(\text{s})$	-3.04

- (a) Describe a standard hydrogen electrode.

(4)

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



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.

(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.

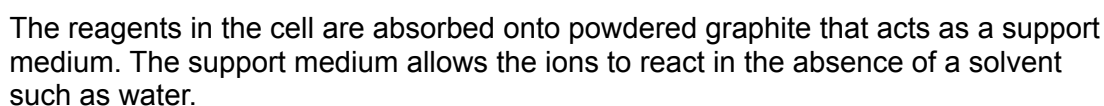
(2)

- (d) A solution of iron(II) sulfate was prepared by dissolving 10.00 g of $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$ ($M_r = 277.9$) in water and making up to 250 cm³ 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 cm³ sample of the partially oxidised solution required 23.70 cm³ of 0.0100 mol dm⁻³ potassium dichromate(VI) solution for complete reaction in the presence of an excess of dilute sulfuric acid.

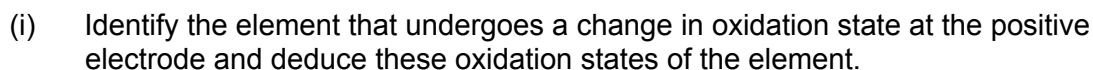
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(Total 14 marks)

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



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



(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 _____

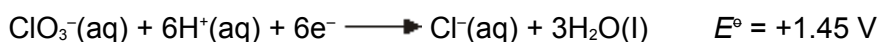
Property 2 _____

(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.



- (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 _____

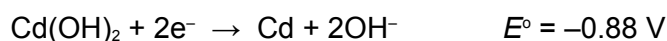
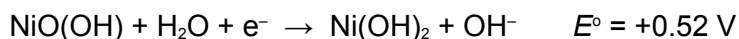
Reducing agent _____

(3)

(Total 12 marks)

Q13.

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.

(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)

Q14.

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° / V
$\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l})$	+1.23
$\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^- \rightarrow 4\text{OH}^-(\text{aq})$	+0.40
$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	0.00
$2\text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightarrow 2\text{OH}^-(\text{aq}) + \text{H}_2(\text{g})$	− 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.

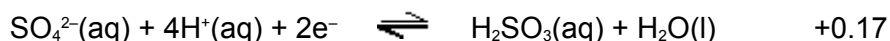
(1)

(Total 10 marks)

Q15.

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

				$E^\ominus / \text{V}$
	$\text{Ce}^{4+}(\text{aq}) + \text{e}^-$	$\rightleftharpoons$	$\text{Ce}^{3+}(\text{aq})$	+1.70
	$\text{MnO}^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^-$	$\rightleftharpoons$	$\text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+1.51
	$\text{Cl}_2(\text{g}) + 2\text{e}^-$	$\rightleftharpoons$	$2\text{Cl}^-(\text{aq})$	+1.36
	$\text{VO}_2^+(\text{aq}) + 2\text{H}^+(\text{aq}) + \text{e}^-$	$\rightleftharpoons$	$\text{VO}^{2+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$	+1.00
	$\text{Fe}^{3+}(\text{aq}) + \text{e}^-$	$\rightleftharpoons$	$\text{Fe}^{2+}(\text{aq})$	+0.77



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

(1)

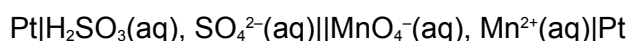
- (b) When the standard electrode potential for $\text{Fe}^{3+}(\text{aq}) / \text{Fe}^{2+}(\text{aq})$ is measured, a platinum electrode is required.

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

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

(3)

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



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 $\text{Fe}^{2+}(\text{aq})$ into $\text{Fe}^{3+}(\text{aq})$ but could not convert $\text{Mn}^{2+}(\text{aq})$ into $\text{MnO}_4^-(\text{aq})$?

(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)

Mark schemes

Q1.

- (a) The ions in the ionic substance in the salt bridge move through the salt bridge

1

To maintain charge balance / complete the circuit

1

- (b) F^-

1

- (c) $E^\ominus SO_4^{2-} / SO_2 < E^\ominus Br_2 / Br^-$

Allow correct answer expressed in words, eg electrode potential for sulfate ions / sulfur dioxide is less than that for bromine / bromide

1

- (d) 1.23 (V)

1

- (e) A fuel cell converts more of the available energy from combustion of hydrogen into kinetic energy of the car / an internal combustion engine wastes more (heat) energy

1

[6]

Q2.

- (a) (i) M1 Positive electrode $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

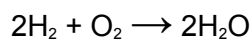
M2 Negative electrode $H_2 + 2OH^- \rightarrow 2H_2O + 2e^-$

Allow multiples, ignore state symbols

If equations both correct but at the wrong electrodes allow 1 mark

1

1



Mark independently

Must be this way round

1

- (ii) Increase (emf)

If decrease/no change then CE=0/2; if blank then mark on

1

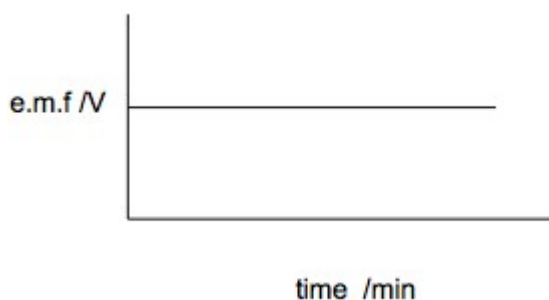


Or overall equation moves to the right

Allow $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$ will move to the right / oxygen half equation moves to the RHS / $E^\ominus O_2/OH^-$ half cell

moves to the right

1



Must start at y-axis

1

- (c) (i) Unchanged

1

- (ii) Water is the only product / fuel cell does not give out pollutants such as NO_x or CO_2 or SO_2 or C or CO or C_xH_y or unburnt hydrocarbons

Not fuel cell does not give out pollutants unless pollutant stated

1

- (d) CO_2 is released because fossil fuels are burned to produce electricity to generate hydrogen
OR

CO_2 is released when methane reacts with steam to produce hydrogen

Allow CO_2 is released to produce the hydrogen

1

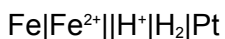
[9]

Q3.

- (a) $\text{Pt}|\text{H}_2|\text{H}^+||\text{Fe}^{2+}|\text{Fe}$

Allow 1 for correct order of symbols but lose second mark for a wrong phase boundary(s) / Pt missing / extra Pt on RHS, additional phase boundary

Note, allow one mark only for correct symbol in reverse:



Allow dashed lines for salt bridge

Ignore state symbols

Ignore 2 if used before H^+

2

- (b) Electron donor

Allow (species that) loses electrons

Do not allow reference to electron pairs

1

- (c) Cl_2 / chlorine

If M1 blank or incorrect cannot score M2

1

(Species on RHS / electron donor) has most positive / largest $E^\ominus$ / has highest potential

Do not allow reference to e.m.f. or $E(\text{cell})$

1

(d) (i) Cl / chlorine

1

(ii) Chlorine +1 to chlorine 0

CE if chlorine not identified in part (i)

Allow chlorine +1 to chlorine -1 (in Cl^-)

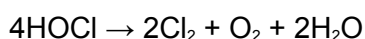
Allow oxidation state decreases by one OR two

Allow oxidation state changes by -1 OR -2

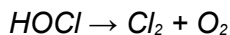
1

(e) $4\text{HOCl} + 4\text{H}^+ + 4\text{OH}^- \rightarrow 2\text{Cl}_2 + \text{O}_2 + 6\text{H}_2\text{O}$

OR



Allow one mark for any incorrect equation that shows



Allow multiples

Ignore state symbols

Penalise one mark for uncanceled or uncombined species

(eg $\text{H}_2\text{O} + \text{H}_2\text{O}$ instead of $2\text{H}_2\text{O}$)

2

(f) (i) e.m.f. = $0.40 - (-1.25) = 1.65 \text{ (V)}$ / $+1.65 \text{ (V)}$

Allow -1.65 (V)

1

(ii) $2\text{Zn} + \text{O}_2 \rightarrow 2\text{ZnO}$

Allow multiples

Ignore state symbols

Do not allow uncanceled species

If more than one equation given, choose the best

1

(iii) A / stainless lid

If M1 incorrect or blank CE=0

1

O_2 (electrode) has a more positive $E^\ominus$ / oxygen (electrode) requires / gains electrons from external circuit

Or reference to the overall equation and a link to electrons going into A

Allow oxygen is reduced and reduction occurs at the positive electrode

OR Zinc (electrode) has more negative $E^\ominus$

Do not allow reference to e.m.f. or $E(\text{cell})$

1

- (iv) (Cell) reaction(s) cannot be reversed / zinc oxide cannot be reduced to zinc by passing a current through it / zinc cannot be regenerated
Allow danger from production of gas / oxygen produced / hydrogen produced

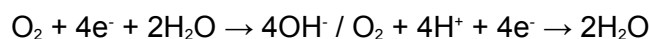
1

[14]

Q4.

- (a) (i) $\text{H}_2 + 2\text{OH}^- \rightarrow 2\text{H}_2\text{O} + 2\text{e}^-$ / $\text{H}_2 \rightarrow 2\text{H}^+ + 2\text{e}^-$
Any order

1



1

- (ii) Hydrogen (electrode) produces electrons
Ignore reference to salt bridge
Do not allow at negative / positive electrode – must identify hydrogen and oxygen

1

Oxygen (electrode) accepts electrons
Allow electrons flow to the oxygen electrode

1

- (b) Hydrogen / the fuel / reactants supplied continuously / fed in
Do not accept oxygen supplied as the only statement

1

- (c) In the fuel cell, a greater proportion of the energy available from the hydrogen–oxygen reaction is converted into useful energy
Allow less energy wasted / more efficient
Do not allow reference to safety

1

- (d) Hydrogen is flammable / H^+ corrosive / OH^- corrosive / hydrogen explosive

1

[7]

Q5.

- (a) Solar cells do not supply electrical energy all the time

1

Rechargeable cells can store electrical energy for use when the solar cells are not working

1

- (b) Prevent pollution of the environment by toxic or dangerous substances / recycling of valuable components
Do not allow 'will not use up landfill sites'.

1

[3]

Q6.

- (a) It has mobile ions / ions can move through it / free ions
Do not allow movement of electrons.
Allow specific ions provided they are moving but do not react.

1

- (b) Chloride ions react with copper ions / Cu^{2+} **OR** $[\text{CuCl}_4]^{2-}$ formed
If incorrect chemistry, mark = 0

1

- (c) The Cu^{2+} ions / CuSO_4 in the left-hand electrode more concentrated
Allow converse.

1

So the reaction of Cu^{2+} with 2e^- will occur (in preference at) left-hand electrode
 $\text{Cu} \rightarrow \text{Cu}^{2+} + \text{electrons}$ at right-hand electrode

Allow left-hand electrode positive / right-hand electrode negative.

Also reduction at left-hand electrode / oxidation at right-hand electrode.

Also left-hand electrode has oxidising agent / right-hand electrode has reducing agent.

Allow E left-hand side $>$ E right-hand side

1

- (d) (Eventually) the copper ions / CuSO_4 in each electrode will be at the same concentration

1

- (e) (i) -3.05 (V)
Must have minus sign.
 -3.05 only.

1

- (ii) $\text{LiMnO}_2 \rightarrow \text{Li} + \text{MnO}_2$ correct equation
Allow 1 for reverse equation.
Allow multiples.

1

Correct direction

If Li^+ not cancelled but otherwise correct, max = 1

If electrons not cancelled, CE = 0

$\text{LiMnO}_2 \rightarrow \text{Li} + \text{MnO}_2$ scores 2

$\text{Li}^+ + \text{LiMnO}_2 \rightarrow \text{Li}^+ + \text{Li} + \text{MnO}_2$ scores 1

$\text{Li} + \text{MnO}_2 \rightarrow \text{LiMnO}_2$ scores 1

1

- (iii) Electricity for recharging the cell may come from power stations burning (fossil) fuel

Allow any reference to burning (of carbon-containing) fuels.

Note combustion = burning.

1

[9]

Q7.

- (a) To remove the oxide layer on the aluminium

Do not allow 'cleaning' or 'removal of grease'.

Do not allow 'removal of impurities' without qualification.

1

- (b) An appropriate method for delivering H₂ gas over a Pt electrode

Need H₂ gas and Pt electrode labelled (allow gas delivered directly below the electrode).

1

The Pt electrode must clearly be in contact with a solution of a named acid.

Ignore any concentration or pressure values.

Ignore absence of bubbles.

Allow if electrode is below outer acid level.

1

- (c) The carbonate ion reacts with the acid (in the SHE) / reaction between carbonate and Al³⁺

Lose this mark if aluminium carbonate formed but mark on.

1

Reaction given (either equation or products specified)

OR H⁺ / Al³⁺ concentrations change / cell e.m.f. altered

1

[5]

Q8.

- (a) Diagram of an Fe³⁺ / Fe²⁺ electrode that includes the following parts labelled:
Solution containing Fe²⁺ and Fe³⁺ ions

1

Platinum electrode connected to one terminal of a voltmeter

Must be in the solution of iron ions (one type will suffice)

1

Salt bridge

Do not allow incorrect material for salt bridge and salt bridge must be in the solution (ie it must be shown crossing a meniscus)

1

298 K and 100 kPa / 1 bar

1

all solutions unit / 1 mol dm⁻³ concentration

Allow zero current / high resistance voltmeter as alternative to M4 or M5

Ignore hydrogen electrode even if incorrect

1

- (b) Cu²⁺ + Fe → Cu + Fe²⁺

Ignore state symbols

1

$\text{Fe}|\text{Fe}^{2+}||\text{Cu}^{2+}|\text{Cu}$ correct order

Allow $\text{Cu}|\text{Cu}^{2+}||\text{Fe}^{2+}|\text{Fe}$

1

Phase boundaries and salt bridge correct, no Pt

Allow single / double dashed line for salt bridge

Penalise phase boundary at either electrode end

Can only score M3 if M2 correct

1

Copper electrode

Allow any reference to copper

1

(c) $E^\ominus \text{Au}^+ / \text{Au} > E^\ominus \text{O}_2 / \text{H}_2\text{O}$

Allow $E_{\text{cell}} / \text{e.m.f.} = 0.45 \text{ V}$

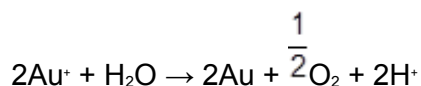
Allow $1.68 > 1.23$

1

So Au^+ ions will oxidise water / water reduces Au^+

QoL

1



Allow multiples

1

(d) $E^\ominus \text{Ag}^+ / \text{Ag} > E^\ominus \text{Fe}^{2+} / \text{Fe}$

Allow $E_{\text{cell}} / \text{e.m.f.} = 1.24$

Allow $0.80 > -0.44$

1

And $E^\ominus \text{Ag}^+ / \text{Ag} > E^\ominus \text{Fe}^{3+} / \text{Fe}^{2+}$

Allow $E_{\text{cell}} / \text{e.m.f.} = 0.03$

Allow $0.80 > 0.77$

1

So silver ions will oxidise iron (to iron(II) ions) and then oxidise Fe(II) ions (further to Fe(III) ions producing silver metal)

Allow Ag^+ ions will oxidise iron to iron(III)

1

[15]

Q9.

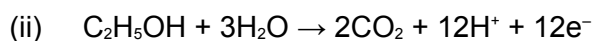
(a) Electron acceptor / gains electrons / takes electrons away

Do not allow electron pair acceptor / gain of electrons / definition of redox (QWC)

1

- (b) $\text{Cd}(\text{OH})_2$
Do not allow 'Cd(OH)₂/Cd'
 1
- Species (on LHS) with the least positive/most negative electrode potential / lowest E / smallest E
Only allow this mark if M1 answer given correctly or blank
Do not allow negative emf
 1
- (c) (i) 1.5 (V) / 1.50
 1
- (ii) $2\text{MnO}_2 + 2\text{H}_2\text{O} + \text{Zn} \rightarrow 2\text{MnO}(\text{OH}) + 2\text{OH}^- + \text{Zn}^{2+}$
Ignore state symbols
 e^- must be cancelled
(take care that Zn^{2+} is on RHS)
 1
- (iii) Allows ions to pass (through it) or words to that effect
Penalise passage of electrons
Allow mention of particular ions
 1
- (iv) Allows electrons to flow / makes electrical contact / conductor
Allow acts as an (inert) electrode / anode / cathode
 1
- (v) Zn is 'used up' / has reacted / oxidised
Allow idea that zinc reacts
Do not allow just zinc corrodes
 1
- (d) (i) 3 / +3 / III
 1
- $2\text{Ni}(\text{OH})_2 + \text{Cd}(\text{OH})_2 \rightarrow 2\text{NiO}(\text{OH}) + \text{Cd} + 2\text{H}_2\text{O}$
For correct nickel and cadmium species in correct order
(allow H_2O missing and OH^- not cancelled)
 1
- For balanced equation (also scores M2)*
Allow max 1 for M2 and M3 if correct balanced equation but reversed.
Ignore state symbols
 1
- (ii) Metal / metal compounds are re-used / supplies are not depleted / It (the cell) can be re-used
Allow does not leak / no landfill problems / less mining / less energy to extract metals / less waste
Do not allow less CO_2 unless explained
 1
- (e) (i) $\text{C}_2\text{H}_5\text{OH} + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O}$
Allow $\text{C}_2\text{H}_6\text{O}$

1



Allow $\text{C}_2\text{H}_6\text{O}$

1

(iii) (+)0.23 (V)

1

(iv) CO_2 released by combustion / fermentation / fuel cell / reaction with water

Can be answered with the aid of equations

1

(atmospheric) CO_2 taken up in photosynthesis

1

[17]

Q10.

(a) loses electrons / donates electrons

penalise donates electron pair

1

(b) Zn

1

(most) negative E° / lowest E° / least positive

can only score M2 if M1 correct

do not allow e.m.f instead of E°

1

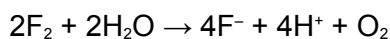
(c) $E^\circ \text{F}_2$ (/F⁻) > $E^\circ \text{O}_2$ (/H₂O)

or e.m.f is positive or e.m.f = 1.64 V

1

Fluorine reacts to form oxygen (can score from equation in M3 even if equation unbalanced provided no contradiction)
or fluorine oxidises water
or fluorine is a more powerful oxidising agent than oxygen

1



allow 4HF in equation

balanced equation scores M2 and M3

1

(d) (i) order correct Zn Zn²⁺ Ag₂O Ag or reverse of this order

ignore ss, H⁺ and H₂O, no. of moles

1

all phase boundaries correct

allow Zn|Zn²⁺||Ag₂O,Ag

or Zn|Zn²⁺||Ag₂O|H⁺|Ag for M1 & M2

e.g. Zn|Zn²⁺||Ag₂O|Ag or Ag|Ag₂O||Zn²⁺|Zn scores 2

M2 cannot be gained unless M1 scored

- allow H^+ either side of Ag_2O with comma or |
for M2 penalise*
- wrong phase boundary (allow dashed lines for salt bridge)*
 - Pt*
 - use of + (from half equation)*
 - water/ H^+ outside Ag in Ag electrode*
- 1
- (ii) 1.1 (V)
- Allow no units, penalise wrong units
allow correct answer even if no answer to (d)(i) or answer to (d)(i) incorrect
allow -1.1 if silver electrode on Left in (d)(i) even if the species are in the wrong order.*
- 1
- (iii) Reaction(s) not reversible or H_2O electrolyses
- do not allow hard to reverse
mention of primary cell is not enough to show that reaction(s) are irreversible*
- 1
- (e) (i) -0.46 (V)
- Allow no units, penalise wrong units*
- 1
- (ii) $2PbSO_4 + 2H_2O \rightarrow Pb + PbO_2 + 2HSO_4^- + 2H^+$
- lead species correct on correct sides of equation
- 1
- equation balanced and includes H_2O ,
 HSO_4^- and H^+ (or H_2SO_4)
- allow ions / species must be fully cancelled out or combined
allow 1/2 for balanced reverse equation*
- 1
- (f) (i) reagents / PbO_2 / H_2SO_4 / acid / ions used up
(or concentration decreases)
- 1
- (ii) fuel cell
- Ignore any other words*
- 1
- (iii) reagents / fuel supplied continuously
- 1
- concentrations (of reagents) remain constant
- 1

[17]

Q11.

(a)	Hydrogen/H ₂ <u>gas/bubbles</u>	1
	1.0 mol dm ⁻³ HCl/H ⁺	1
	At 298K and 100kPa <i>Allow 1 bar instead of 100 kPa</i> <i>Do not allow 1 atm</i>	1
	Pt (electrode)	1
(b)	Li ⁺ + MnO ₂ + e ⁻ → LiMnO ₂ <i>Ignore state symbols</i>	1
	-0.13(V)	1
(c)	Fe ³⁺ ions reduced to Fe ²⁺ <i>Can score from equation/scheme</i>	1
	Because $E(\text{Fe}^{3+}/\text{Fe}^{2+}) > E(\text{H}^+/\text{H}_2)/E(\text{hydrogen})$ <i>Allow emf/E_{cell} +ve/0.77V</i> <i>Allow Fe³⁺ better oxidising agent than H⁺</i> <i>Allow H₂ better reducing agent than Fe²⁺</i> <i>Only award this explanation mark if previous mark given</i>	1
(d)	Moles Cr ₂ O ₇ ²⁻ = <u>23.7 × 0.01/1000</u> = 2.37 × 10 ⁻⁴	1
	1 mol Cr ₂ O ₇ ²⁻ reacts with 6 mol Fe ²⁺ so moles Fe ²⁺ in 25 cm ³ = 6 × 2.37 × 10 ⁻⁴ = 1.422 × 10 ⁻³	1
	$M1 \times 6$	
	Moles Fe ²⁺ in 250 cm ³ = 1.422 × 10 ⁻²	
	$M2 \times 10$ or $M4/10$	1
	Original moles Fe ²⁺ = <u>10.00/277.9</u> = 0.0360 <i>Independent mark</i>	1
	Moles Fe ²⁺ oxidised = 0.0360 – 0.0142 = 0.0218 $M4 - M3$	1
	% oxidised = (0.0218 × 100)/0.0360 = 60.5% $(M5 \times 100)/M4$	

Allow 60 to 61
 Note Max 3 if mol ratio for M2 wrong
 eg 1:5 gives 67.1%
 1:1 gives 93.4%
 Note also, 39.5% (39-40) scores M1, M2, M3 and M4 (4 marks)

1

[14]

Q12.

- (a) (i) Co/Cobalt

If Co or Cobalt not given CE = 0
 ignore case in symbol for Co

1

(+) 4

1

(+) 3

Allow 4 and 3 in either order

1

- (ii) $\text{Li} \rightarrow \text{Li}^+ + \text{e}^-$

Ignore state symbols
 Allow e without -ve sign
 Do not allow equilibrium sign

1

- (iii) Platinum is a conductor

1

(Platinum is) unreactive/inert

Ignore mention of surface area or catalyst
 Allow 2 marks if two properties given on one answer line
 Apply list principle to contradictions/wrong answers
 Do not allow platinum resists corrosion

1

- (iv) Li reacts with water/forms lithium hydroxide

Allow water breaks down (or is electrolysed) on re-charge

1

- (b) (i) $\text{Pt} | \text{SO}_3^{2-}(\text{aq}), \text{SO}_4^{2-}(\text{aq}) || \text{ClO}_3^-(\text{aq}), \text{Cl}^-(\text{aq}) | \text{Pt}$

State symbols an ', ' not necessary
 Allow | in place of ', ' NOT ', ' in place of |
 Ignore H^+ and H_2O
 Deduct one mark for each mistake (e.g. Pt missed twice counts as two mistakes)
 Allow reverse order for whole cell
 $\text{Pt} | \text{Cl}^-, \text{ClO}_3^- || \text{SO}_4^{2-}, \text{SO}_3^{2-} | \text{Pt}$

2

- (ii) $\text{ClO}_3^- + 3\text{SO}_3^{2-} \rightarrow \text{Cl}^- + 3\text{SO}_4^{2-}$

Oxidising agent ClO_3^-

Reducing agent SO_3^{2-}

[12]

Q13.

(a) 1.4 V

Allow + or –

(b) $2\text{NiO}(\text{OH}) + 2\text{H}_2\text{O} + \text{Cd} \rightarrow 2\text{Ni}(\text{OH})_2 + \text{Cd}(\text{OH})_2$

*Mark for species, Deduct a mark for additional species
(eg OH-) but allow balance mark*

Balanced

If equation is reversed CE=0

(c) $\text{NiO}(\text{OH})$ or $\text{Ni}(\text{III})$ or nickel

+3

Allow conseq on wrong species

[5]

Q14.

(a) By definition

allow 'set to this value'

(b) 1.23 V

Allow + or –

(c) $\text{Pt}|\text{H}_2(\text{g})|\text{OH}^-(\text{aq}), \text{H}_2\text{O}(\text{l})||\text{O}_2(\text{g})|\text{H}_2\text{O}(\text{l}), \text{OH}^-(\text{aq})|\text{Pt}$

H₂O not essential, allow reverse order

Correct but with Pt missing

Includes Pt with correct representation

(d) Uses $\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow 4\text{OH}^-$

And (2×) $2\text{OH}^- + \text{H}_2 \rightarrow 2\text{H}_2\text{O} + 2\text{e}^-$

- $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ 1
- (e) Increases the surface area (so reaction faster) 1
- (f) Overall reaction is the same ($2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$)
Or shows e.m.f. is the same 1
- (g) Hydrogen and oxygen supplied continuously
OR
Can be operated without stopping to recharge
Or can be refuelled quickly
Allow any one mark 1
- (h) Hydrogen may need to be made using an energy source that is not 'carbon neutral' 1

[10]

Q15.

- (a) (Standard) hydrogen (electrode) (1) 1
- (b) (i) To allow transfer of electrons / provide a reaction surface (1)
- (ii) 298 K (1)
Both $\text{F}^{3+}(\text{aq})$ and $\text{Fe}^{2+}(\text{aq})$ have a concentration of
 1 mol dm^{-3} (1) (QoL)
OR $[\text{H}^+] = 1 \text{ mol dm}^{-3}$
NOT zero current or 100 kPa 3
- (c) +1.34 V (1)
 $2 \text{MnO}_4^- + 5 \text{H}_2\text{SO}_3 \rightarrow 2 \text{Mn}^{2+} + 5 \text{SO}_4^{2-} + 3 \text{H}_2\text{O} + 4 \text{H}^+$
Correct species / order (1)
Balanced and cancelled (1)
Allow one for $2 \text{MnO}_4^- + 5 \text{H}_2\text{SO}_3 \rightarrow 2 \text{Mn}^{2+} + 5 \text{SO}_4^{2-}$ 3
- (d) (i) $\text{Ce}^{4+}(\text{aq})$ (1)
- (ii) $\text{VO}_2^+(\text{aq})$ (1); Cl_2 (1)
Penalise additional answers to zero 3
- (e) $\text{Pt} | \text{Fe}^{2+}(\text{aq}), \text{Fe}^{3+}(\text{aq}) || \text{Ce}^{4+}(\text{aq}), \text{Ce}^{3+}(\text{aq}) | \text{Pt}$
Correct species (1)
Correct order (1)
Deduct one mark for each error 2

[12]

Examiner reports

Q2.

Only the more able students scored well in part (a)(i). Many students failed to give the correct equations at the electrodes and the overall equation was often not cancelled down to the simplest form. Most students scored the first mark in part (a)(ii) but only a very few students referred to an equation to score M2. Most students could draw the correct line on the graph in part (b) although several assorted incorrect curves were seen. The effect of changing the surface area was well known in part (c)(i). Some students failed to gain credit in part (d) because their answers were vague and did not identify pollutants. In part (e) a statement about carbon dioxide being released in the generation of hydrogen was required to score the mark; some students did not identify hydrogen as the fuel in the fuel cell and described general ways that carbon dioxide would be released when the fuel-cell was made e.g. transportation of hydrogen etc.

Q3.

Few correct answers were seen to part (a). Some students did not comply with the convention of placing the hydrogen electrode on the left when measuring the standard electrode potential. Some failed to include all the phase boundaries, putting a comma between H_2 and H^+ instead of a solidus, whilst other students included too many phase boundaries at the extremities or an incorrect, additional platinum connected to iron. Parts (b) and (c) were answered well. All the elements were chosen by students in part (d)(i) and where the correct answer of chlorine was selected, the oxidation states in part (d)(ii) and the equation in part (e) were usually correct. It was pleasing to find that answers to parts (f)(i) and (f)(ii) were usually correct. Most students identified the positive electrode correctly in part (f)(iii) but only the highest-scoring students used an appropriate electrode potential from the table to explain their answer.

Q4.

Answers to parts of this question were generally disappointing. In part (a)(i), the half-equation for the hydrogen electrode was usually correct but most students gave an incorrect equation for the reaction at the oxygen electrode involving the formation of oxide ions. Part (a)(ii) proved to be difficult and only the highest-scoring students were able to address the question properly by using the half equations to explain how a flow of electrons is generated in the external circuit.

Part (b) was answered well but answers to part (c) were mostly disappointing. Students appeared not to know that the major advantage of a fuel cell over a heat engine is its greater efficiency. It is difficult for examiners to understand why most students thought that the burning of hydrogen in an internal combustion engine would lead to formation of carbon dioxide and hence exacerbate global warming.

Q5.

In this part (a), there were many examples of full marks being awarded for vague, unclear or confusing answers. Unless the student makes it very clear which type of cell they are referring to in their answer they cannot gain full credit. Lenient marking was often seen here. The ideas in this part (b), however, were well understood.

Despite the long catalogue of points listed above, this was, overall, another very positive session in many schools and colleges. The new administrative arrangements were applied well and schools and colleges are thanked for this. Teachers in schools and

colleges have, in the main, successfully and securely delivered the practical assessment and have impressed with the accurate and consistent standard of their marking. In my first report as Principal Moderator, I am grateful for the opportunity to salute their professionalism and thank them for their integrity.

Q6.

This was a testing question and in common with redox questions in previous years, performance by students was very mixed. In part (a), over 30% of students answered incorrectly that the salt bridge allowed electrons to flow through it. In part (b), students were expected to use their knowledge of copper chemistry to suggest that the chloride ions from the potassium chloride would react with the aqueous copper ions but only a very small percentage of students answered this correctly. Part (c) also proved to be a very demanding question. Only the best students gave a correct answer involving the higher concentration of copper ions in the left hand electrode. It was surprising that many students appeared to consider that the 0.2 mol dm^{-3} solution of CuSO_4 was more concentrated than the 1.0 mol dm^{-3} solution. Part (d) was very discriminating. The great majority of students answered incorrectly by stating that one of the reagents would be used up or had reacted completely. Answers to part (e) were much better. When students are asked to write an equation for an overall reaction, they should take care to cancel out identical species that appear on both sides of the final equation. In part (e)(ii), many students lost a mark because they failed to cancel out a lithium ion that appeared on both sides of the equation. In part (e)(iii), it was surprising to note that a considerable number of students thought that carbon dioxide would be produced by a reaction with an imaginary carbon electrode in the lithium cell.

Q7.

- (a) This part was usually well answered. The usual mistakes were to make some reference to impurities without specifying aluminium oxide or contradicting an answer by giving an incorrect formula.
- (b) Answers to this part were variable. Students familiar with the standard hydrogen electrode gave good answers, although many omitted to name the acid used.
- (c) In this part, most students realised that a reaction would occur but gave impossible products and lost the mark. The consequences of reaction were not well understood. A number of students gave answers implying that a salt bridge would not be necessary or describing details of one specific use of a standard hydrogen electrode, which did not relate to the question being asked.

Q8.

This proved to be a difficult question. In part (a), answers were often very scrappy and disappointing. Diagrams showed a half cell that had no chance of working. The salt bridge, if present rarely dipped into a solution and the electrode material was often iron instead of the correct material, platinum. Only very good students showed a solution containing iron(II) and iron(III) ions and only the very highest-scoring students stated clearly that all aqueous species should have a concentration of 1 mol dm^{-3} . Answers to part (b) were more encouraging but only the better students realised that no platinum electrode are necessary with copper and iron electrodes. In parts (c) and (d), most students were unable to use data from the table correctly. For instance, it is correct to state that $E^\ominus \text{Au}^+/\text{Au} > E^\ominus \text{O}_2/\text{H}_2\text{O}$ but not correct to state that $E^\ominus \text{Au}^+/\text{Au} > E^\ominus \text{H}_2\text{O}/\text{O}_2$ or that gold ions are better oxidising agents than water. In this reaction, water is the reducing agent. In part (d), only 6% of students recognised that silver ions would oxidise iron to iron(II) ions and then go on to oxidise iron(II) ions to iron(III) ions.

Q9.

Part (a) was answered well. The most common error was to give the meaning of oxidation rather than an oxidising agent. Answers to part (b) were very often incorrect. All of the species from the table were seen in students' answers. The most common wrong answer was probably 'cadmium'. It still seems to be the case that many students do not understand that when redox half-equations are presented in their standard form with electrons on the left, the species on the left are oxidising agents and the species on the right are reducing agents. Parts (c)(i) and (ii) were answered correctly by most students. Part (c)(iii) proved to be more testing and only the best students recognised that the porous partition was acting in the same way as a salt bridge by permitting the transport of ions. Parts (c)(iv) and (v) were answered quite well although in part (c)(v) it was not always realised that the zinc container reacts to form zinc ions whilst the cell is in operation.

Part (d)(i) proved to be a discriminating question. An oxidation state of +2 for nickel in $\text{NiO}(\text{OH})$ was a common error and the equation was often unbalanced or not simplified by removing hydroxide ions. Part (d)(ii) was almost always answered correctly. In comparison, parts (e)(i) and (ii) were not answered well, usually because of a lack of attention to detail. Instead of using ethanol, a common mistake was to use CH_3OH . In part (e)(i), the equation for the combustion of ethanol was often unbalanced because students seemed to fail to take into account the oxygen in the ethanol. In part (e)(ii), despite being asked to write a half-equation, the electrons were often omitted or the numbers of hydrogen ions and electrons were wrong. Answers to part (e)(iii) were usually correct. In part (e)(iv), it seems that many students are still not clear about the meaning of the term carbon-neutral. The term is too often thought to refer just to carbon rather than to carbon dioxide in the atmosphere. Also, the majority of students did not make it clear that the carbon dioxide, released into the atmosphere by the processes that manufacture and use ethanol, can be balanced by the carbon dioxide taken out of the atmosphere by photosynthesis in plants.

Q10.

As usual, this question that tested candidates' understanding of redox and cells proved to be difficult but it was pleasing to note that even the weaker candidates were able to gain a significant number of marks. Parts (a) and (b) were answered well. Many candidates were able to score two out of three marks for part (c). The mark lost was usually due to incorrect use of data from the table. Many candidates stated that the electrode potential values showed that fluorine is a better oxidising agent than water. In this case, water is the reducing agent and candidates should have stated that the data shows that fluorine is a better oxidising agent than oxygen. In part (d)(i), a majority of candidates scored one mark out of the two available; only the best candidates showed metal species on the outside of the cell and all the phase boundaries, correctly indicated. Parts (d)(ii) and (e)(i) were answered well but only about a third of candidates gave a correct answer to part (d)(iii). In part (e)(ii) marks were lost because candidates did not simplify the overall equation after combining the half equations or the equation was for discharge rather than for recharging. As expected, part (f) proved to be challenging, particularly part (f)(iii). Fuel cells were included in the Specification as an example of 'How Science Works'. Their value is that they can produce efficiently a steady current to drive electric motors because the concentrations of the reactants remain constant when the fuel and the oxidant are supplied continuously.

Q11.

This was the most difficult question in the paper. In part (a) most candidates knew that a standard hydrogen electrode uses hydrogen gas and a platinum electrode but only a minority went on to state the correct and necessary conditions of concentration,

temperature and pressure for operation of the electrode. In part (b) most candidates gave a correct value for the electrode potential of the positive electrode but very few candidates were able to use the cell representation to deduce a half-equation for the electrode reaction. Answers to part (c) were also disappointing. Despite the given electrode potentials and half-equations, most candidates could not predict and explain the redox reaction between hydrogen and iron(III) ions. Part (d) also proved to be a difficult question. Candidates were usually able to score three marks out of six for a correct calculation of the number of moles of dichromate(VI) ions, a calculation of the original number of moles of iron(II) sulfate and for recognising the involvement of a factor of ten. Only the very best candidates realised that the number of moles of iron ions was six times the number of moles of dichromate(VI) ions and that a subtraction was required to calculate the percentage of iron(II) ions that had been oxidised by air before the titration.

Q12.

Redox questions involving electrochemical cells are usually found to be very difficult. It was therefore very pleasing to find that this question was answered well in all parts except for parts (a)(iv) and (b)(i). In part (a)(iv) candidates were expected to recall from Unit 1 that lithium reacts with water but only about 20% were able to do this and these were not necessarily the best candidates. As expected, part (b)(i) proved to be discriminating. Although about one third of candidates scored both marks, others failed to include platinum or put the various ions in the wrong order.

Q13.

This question was answered well except for part (b) where only the best candidates were able to write a balanced equation with the oxidising and reducing agent in the correct position.

Q14.

This question covered a new area of the Specification although the principles of redox that were involved have been tested for many years. Candidates have always found difficulty with the understanding of these principles when applied to electrochemical cells. This question discriminated very effectively between candidates. In part (a) only about 20% of candidates knew that the zero electrode potential of the hydrogen electrode is an arbitrary definition. In part (c) candidates had great difficulty with the conventional representation of the alkaline fuel cell. The most common mistakes were to give the oxidised and reduced forms in the wrong order and to use the hydrogen half-equation in acidic rather than alkaline conditions. Similar mistakes were made in part (d) although 39% of candidates did score full marks. About half of all candidates scored full marks in parts (g) and (h) but only 23% of candidates gained the mark in part (f). Many candidates opted not to answer this part.

Q15.

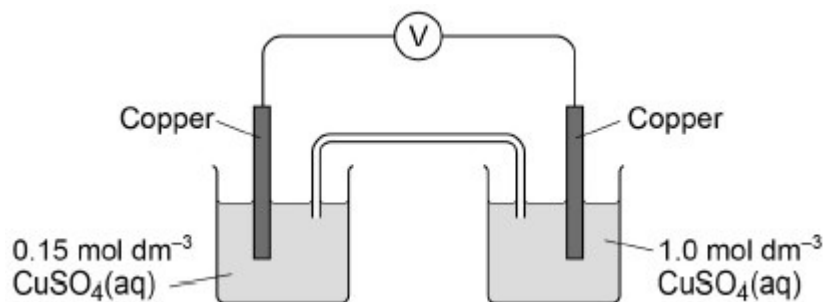
There were many good answers to this question. Virtually all candidates were able to name the standard reference electrode in part (a). In part (b) only the best candidates stated the function of the platinum electrode correctly and many failed to state in section (ii) that, for standard conditions to apply, both $\text{Fe}^{3+}(\text{aq})$ and $\text{Fe}^{2+}(\text{aq})$ would need to be at a concentration of 1 mol dm^{-3} in the solution. Most deduced correctly the e.m.f. of the cell given in part (c).

Many scored one further mark in part (c) when they deduced the correct molar ratio of the reacting species in the equation but lost one mark when they failed to balance the equation correctly. In part (d) most candidates were able to deduce which of the species in the table of data was the strongest oxidising agent but many only gave one of the two

correct answers in section (ii). A pleasing number of candidates gave a correct cell representation in part (e). Almost all were able to select the correct species for the cell but some then used the convention, as shown in part (c), incorrectly.

Q1.

A student set up the cell shown in the diagram.

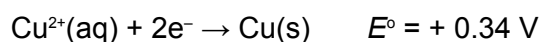


The student recorded an initial voltage of +0.16 V at 25 °C

- (a) Explain how the salt bridge provides an electrical connection between the two solutions.

(1)

- (b) The standard electrode potential for the Cu^{2+}/Cu electrode is



Calculate the electrode potential of the left-hand electrode in the diagram.

Electrode potential _____ V

(1)

- (c) Both electrodes contain a strip of copper metal in a solution of aqueous Cu^{2+} ions.
State why the left-hand electrode does **not** have an electrode potential of +0.34 V

(1)

- (d) Give the conventional representation for the cell in the diagram.
Include all state symbols.

(1)

- (e) When the voltmeter is replaced by a bulb, the EMF of the cell in the diagram decreases over time to 0 V

Suggest how the concentration of copper(II) ions in the left-hand electrode changes when the bulb is alight.

Give **one** reason why the EMF of the cell decreases to 0 V

Change in concentration of copper(II) ions in the left-hand electrode

Reason why the EMF decreases to 0 V _____

(2)
(Total 6 marks)

Q2.

This question is about vanadium compounds and ions.

- (a) Use data from Table 4 to identify the species that can be used to reduce VO_2^+ ions to VO^{2+} in aqueous solution and no further.

Explain your answer.

Electrode half-equation	$E^\ominus / \text{V}$
$\text{VO}_2^+(\text{aq}) + 2\text{H}^+(\text{aq}) + \text{e}^- \rightarrow \text{VO}^{2+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$	+1.00
$\text{VO}^{2+}(\text{aq}) + 2\text{H}^+(\text{aq}) + \text{e}^- \rightarrow \text{V}^{3+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$	+0.34
$\text{Cl}_2(\text{aq}) + 2\text{e}^- \rightarrow 2\text{Cl}^-(\text{aq})$	+1.36
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq})$	+0.77
$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn}(\text{s})$	-0.76

Reagent _____

Justification _____

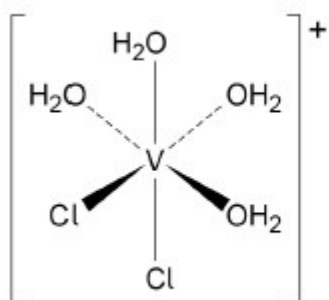
(2)

- (b) Give the oxidation state of vanadium in $[\text{VO}(\text{H}_2\text{O})_5]^{2+}$

(1)

- (c) The $[\text{V}(\text{H}_2\text{O})_4\text{Cl}_2]^+$ ion exists as two isomers. One isomer is shown.

Draw the structure of the other isomer and state the type of isomerism.



Type of isomerism _____

(2)

- (d) Heating NH_4VO_3 produces vanadium(V) oxide, water and one other product.

Give an equation for the reaction.

(1)

- (e) Vanadium(V) oxide is the catalyst used in the manufacture of sulfur trioxide.

Give **two** equations to show how the catalyst is used and regenerated.

(1)

(Total 7 marks)

Q3.

Which ion **cannot** catalyse the reaction between iodide (I^-) and peroxodisulfate ($\text{S}_2\text{O}_8^{2-}$)?

Use the data below to help you answer this question.

Half-equation	E° / V
$\text{S}_2\text{O}_8^{2-} + 2\text{e}^- \rightarrow 2\text{SO}_4^{2-}$	+2.01
$\text{Co}^{3+} + \text{e}^- \rightarrow \text{Co}^{2+}$	+1.82
$\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$	+0.77
$\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$	+0.54
$\text{Cr}^{3+} + \text{e}^- \rightarrow \text{Cr}^{2+}$	-0.41

A Co^{2+}

☐

B Cr^{2+}

☐

C Fe^{2+}

☐

**Q4.**

The table shows some electrode half-equations and their standard electrode potentials.

Electrode half-equation	E° / V
$\text{Cl}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{Cl}^-(\text{aq})$	+1.36
$\text{NO}_3^-(\text{aq}) + 4\text{H}^+(\text{aq}) + 3\text{e}^- \rightarrow \text{NO}(\text{aq}) + 2\text{H}_2\text{O}(\text{aq})$	+0.96
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq})$	+0.77
$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$	+0.34
$\text{SO}_4^{2-}(\text{aq}) + 4\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{SO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{aq})$	+0.17
$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	0.00
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Fe}(\text{s})$	-0.44

- (a) Deduce the oxidation state of nitrogen in NO_3^- and in NO

Nitrogen in NO_3^- _____

Nitrogen in NO _____

(2)

- (b) State the weakest reducing agent in the table.

(1)

- (c) Use data from the table to identify an acid that will oxidise copper.

Explain your choice of acid.

Use these data to suggest a possible equation for the reaction.

Calculate the EMF of the cell that has the same overall reaction.

Q5.

The table shows some standard electrode potential data.

Electrode half-equation	$E^\ominus / \text{V}$
$\text{Cl}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{Cl}^-(\text{aq})$	+1.36
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$	+0.80
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq})$	+0.77
$\text{Sn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Sn}(\text{s})$	-0.14
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Fe}(\text{s})$	-0.44

- (a) Use data from the table to deduce the species that is the best oxidising agent.

(1)

- (b) Write the conventional representation for the cell used to measure the standard electrode potential for the conversion of tin(II) ions to tin.

(2)

- (c) A cell was made by connecting two half-cells with a salt bridge. One half-cell consisted of silver in a solution of silver nitrate solution and the other consisted of tin in a solution of tin(II) nitrate solution.

Calculate the EMF of this cell and write a half-equation for the reaction that occurs at the negative electrode.

EMF _____

Half-equation

(2)

- (d) Use data from the table above to write an equation for the reaction of silver(I) ions with iron(II) ions.

(1)
(Total 6 marks)

Q6.

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	$E^\ominus / \text{V}$
$\text{F}_2 + 2\text{e}^- \longrightarrow 2\text{F}^-$	+2.87
$\text{Cl}_2 + 2\text{e}^- \longrightarrow 2\text{Cl}^-$	+1.36
$\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \longrightarrow 2\text{H}_2\text{O}$	+1.23
$\text{Br}_2 + 2\text{e}^- \longrightarrow 2\text{Br}^-$	+1.07
$\text{I}_2 + 2\text{e}^- \longrightarrow 2\text{I}^-$	+0.54
$\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \longrightarrow 4\text{OH}^-$	+0.40
$\text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^- \longrightarrow \text{SO}_2 + 2\text{H}_2\text{O}$	+0.17
$2\text{H}^+ + 2\text{e}^- \longrightarrow \text{H}_2$	0.00
$4\text{H}_2\text{O} + 4\text{e}^- \longrightarrow 4\text{OH}^- + 2\text{H}_2$	-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.

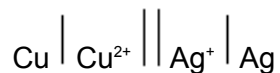
Suggest the major advantage of using the fuel cell.

(1)

(Total 6 marks)

Q7.

The following cell has an EMF of +0.46 V.



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)

Q8.

Iron used in the construction of pipes, bridges and ships, can be protected by attaching a sacrificial anode (electrode) to the structure. This sacrificial anode is another metal that corrodes in preference to iron, keeping the iron structure unaffected. This metal must be a stronger reducing agent than iron.

The table shows the standard electrode potentials for some common metal electrodes in alphabetical order.

	Electrode equation	Standard electrode potential / V
aluminium	$\text{Al}^{3+}(\text{aq}) + 3\text{e}^{-} \rightarrow \text{Al}(\text{s})$	-1.66
copper	$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Cu}(\text{s})$	+0.34
iron	$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Fe}(\text{s})$	-0.44
lead	$\text{Pb}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Pb}(\text{s})$	-0.13
magnesium	$\text{Mg}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Mg}(\text{s})$	-2.37
tin	$\text{Sn}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Sn}(\text{s})$	-0.14
zinc	$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Zn}(\text{s})$	-0.76

- (a) Use the standard electrode potentials in the table to identify the most effective metal for use as a sacrificial anode for iron and explain your answer.

(2)

- (b) The pH of seawater varies between 7.5 and 8.4. An aluminium sacrificial anode slowly corrodes during its use. When attached to a ship it comes into contact with seawater.
The aluminium sets up a cell between itself and the iron structure. In this cell, oxygen dissolved in water is reduced to hydroxide ions on the iron surface. This oxygen/hydroxide electrode has a standard electrode potential of +0.40 V.

Write a half-equation for the oxygen/hydroxide electrode reaction.

Construct an equation to show the overall cell reaction involving the aluminium anode.

Use the table and the information above to calculate the e.m.f. of the cell.

Half-equation _____

Overall cell reaction _____

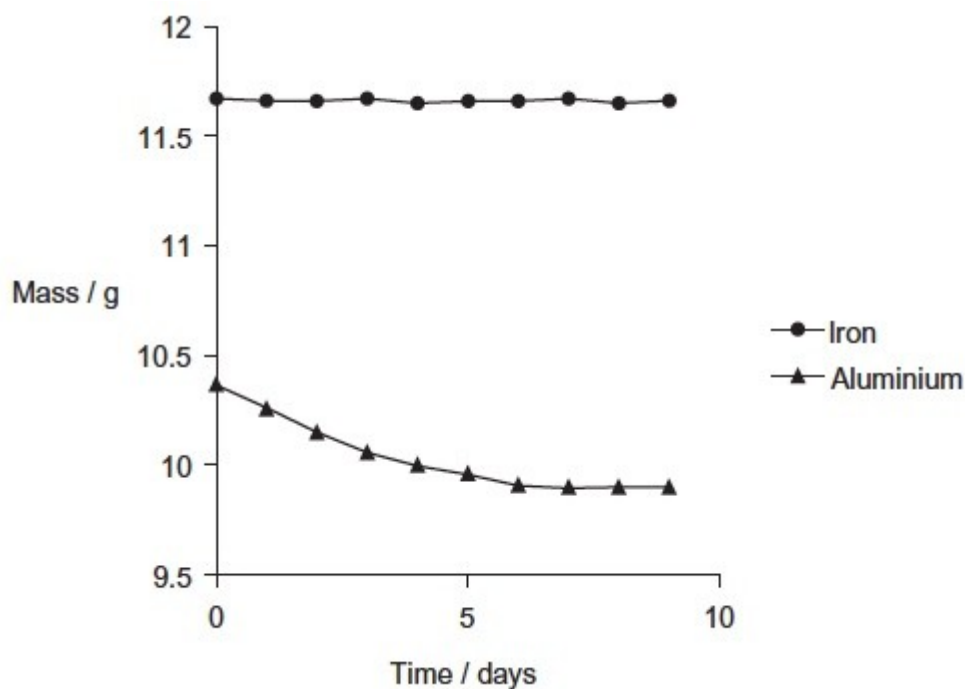
Cell e.m.f. _____

(3)

- (c) Give **two** reasons why the value of the cell e.m.f that you calculated in part (b) would not normally be achieved when the cell operates in seawater.

(2)

To examine the effect of a sacrificial anode, a student carefully cleaned a piece of iron foil and a piece of aluminium foil. These were joined and placed in a beaker containing a buffer solution at pH 8.0. Air was bubbled steadily into the solution. At intervals the foils were removed, separated, dried and weighed before being placed back in the buffer solution. The results obtained were plotted on the graph.



(d) Explain briefly the results obtained for the iron foil.

(1)

(e) Suggest briefly why the aluminium foil stopped decreasing in mass after a time.

(1)

(Total 9 marks)

Q9.

The table shows some standard electrode potential data.

Electrode half-reaction	$E^\ominus/\text{V}$
$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn}(\text{s})$	-0.76
$\text{Co}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Co}(\text{s})$	-0.28
$\frac{1}{2}\text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2\text{O}(\text{l})$	+1.23
$\text{Au}^+(\text{aq}) + \text{e}^- \rightarrow \text{Au}(\text{s})$	+1.68
$\text{Co}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Co}^{2+}(\text{aq})$	+1.82

- (a) (i) Identify the weakest oxidising agent in the table.

(1)

- (ii) Give the conditions under which the electrode potential of the Zn^{2+}/Zn electrode is -0.76 V.

(2)

- (b) Two half-cells, involving species in the table, are connected together to give a cell with an e.m.f. = +0.48 V.

Use data from the table to deduce the conventional representation of this cell. Write the half-equation for the reaction that occurs at the negative electrode.

Conventional representation _____

Half-equation _____

(3)

- (c) Use data from the table to identify a cobalt species that can react with water.

Write an equation for the redox reaction that occurs and identify the oxidation product in the reaction.

Cobalt species _____

Equation _____

Oxidation product _____

(3)

- (d) Use data from the table to explain why gold jewellery is unreactive in moist air.

(2)

(Total 11 marks)

Q10.

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	$E^\ominus / \text{V}$
$\text{F}_2 + 2\text{e}^- \longrightarrow 2\text{F}^-$	+2.87
$\text{Au}^+ + \text{e}^- \longrightarrow \text{Au}$	+1.68
$2\text{HOCl} + 2\text{H}^+ + 2\text{e}^- \longrightarrow \text{Cl}_2 + 2\text{H}_2\text{O}$	+1.64
$\text{Cl}_2 + 2\text{e}^- \longrightarrow 2\text{Cl}^-$	+1.36

$\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \longrightarrow 2\text{H}_2\text{O}$	+1.23
$\text{Ag}^+ + \text{e}^- \longrightarrow \text{Ag}$	+0.80
$\text{Fe}^{3+} + \text{e}^- \longrightarrow \text{Fe}^{2+}$	+0.77
$2\text{H}^+ + 2\text{e}^- \longrightarrow \text{H}_2$	0.00
$\text{Fe}^{2+} + 2\text{e}^- \longrightarrow \text{Fe}$	-0.44

(e) A cell is made by connecting $\text{Fe}^{2+} / \text{Fe}$ 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.

(2)

(Total 15 marks)

Q11.

The table below shows some standard electrode potential data.

	$E^\ominus / \text{V}$
$\text{ZnO(s)} + \text{H}_2\text{O(l)} + 2\text{e}^- \longrightarrow \text{Zn(s)} + 2\text{OH}^-(\text{aq})$	-1.25
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Fe(s)}$	-0.44

$\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^- \longrightarrow 4\text{OH}^-(\text{aq})$	+0.40
$2\text{HOCl}(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{Cl}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$	+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.

(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

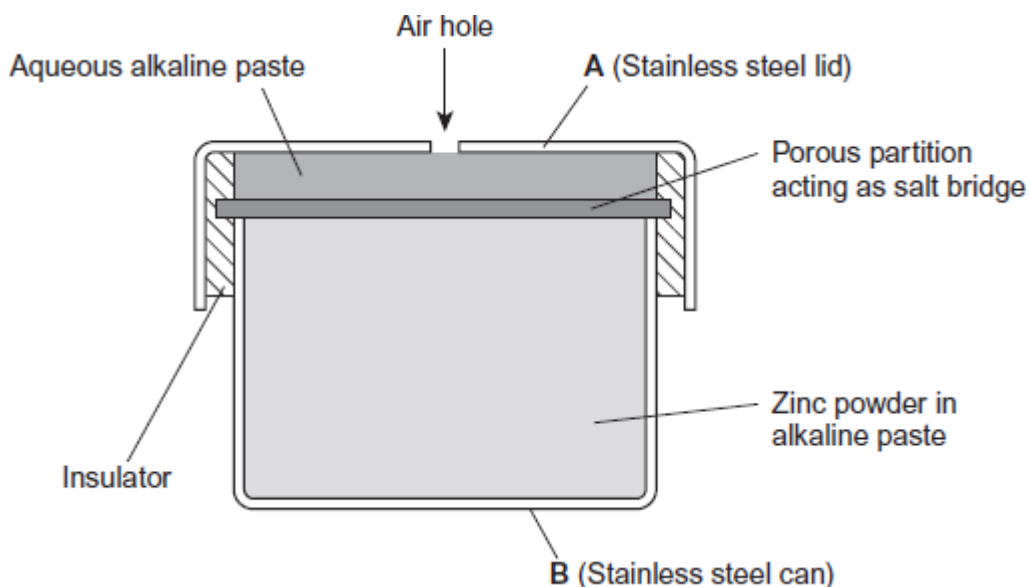
(2)

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

	$E^\ominus / \text{V}$
$\text{ZnO(s)} + \text{H}_2\text{O(l)} + 2\text{e}^- \longrightarrow \text{Zn(s)} + 2\text{OH}^-(\text{aq})$	-1.25
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Fe(s)}$	-0.44
$\text{O}_2(\text{g}) + 2\text{H}_2\text{O(l)} + 4\text{e}^- \longrightarrow 4\text{OH}^-(\text{aq})$	+0.40
$2\text{HOCl(aq)} + 2\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{Cl}_2(\text{g}) + 2\text{H}_2\text{O(l)}$	+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.

- Positive electrode _____

Q12.

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.

- (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)

Q13.

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



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 cm³ 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.

(3)

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

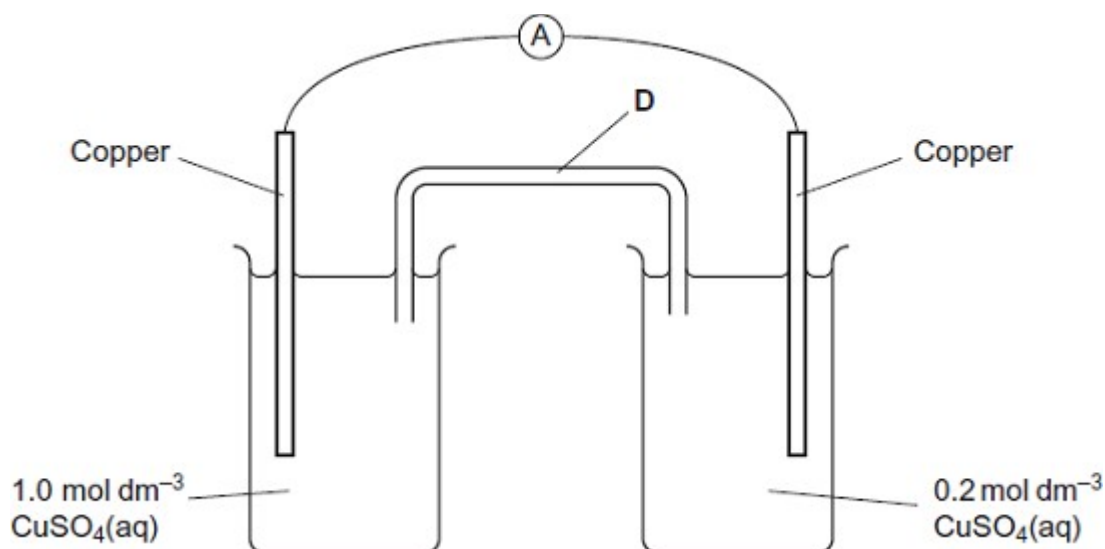
Briefly outline how you would carry out an experiment to confirm your answer to part (b).

(3)

(Total 8 marks)

Q14.

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.

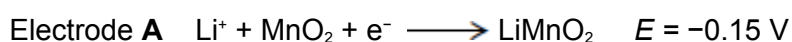
Suggest why the electrons move in this direction.

(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 **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**.

(2)

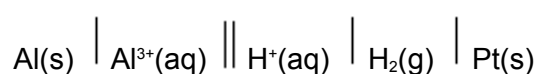
- (iii) Suggest why the recharging of a lithium cell may lead to release of carbon dioxide into the atmosphere.

(1)

(Total 9 marks)

Q15.

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



- (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.

(2)

(Total 5 marks)

Q16.

This table shows some standard electrode potential data.

Electrode half-equation	$E^\ominus / \text{V}$
$\text{Au}^+(\text{aq}) + \text{e}^- \longrightarrow \text{Au}(\text{s})$	+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 _____

(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.

(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.

(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.

(3)

(Total 15 marks)

Q17.

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*.

(1)

- (b) Deduce the identity of the weakest oxidising agent in the table.

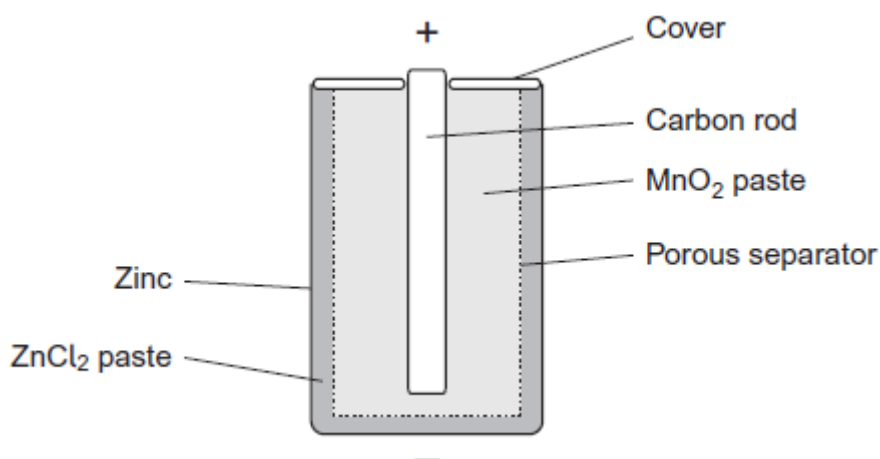
Explain how $E^\ominus$ values can be used to make this deduction.

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.

(2)

(Total 17 marks)

Q18.

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

	$E^\ominus / \text{V}$
$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+1.51
$\text{Cl}_2(\text{aq}) + 2\text{e}^- \rightarrow 2\text{Cl}^-(\text{aq})$	+1.36
$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{aq})$	0.00

(2)

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

(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.

(2)

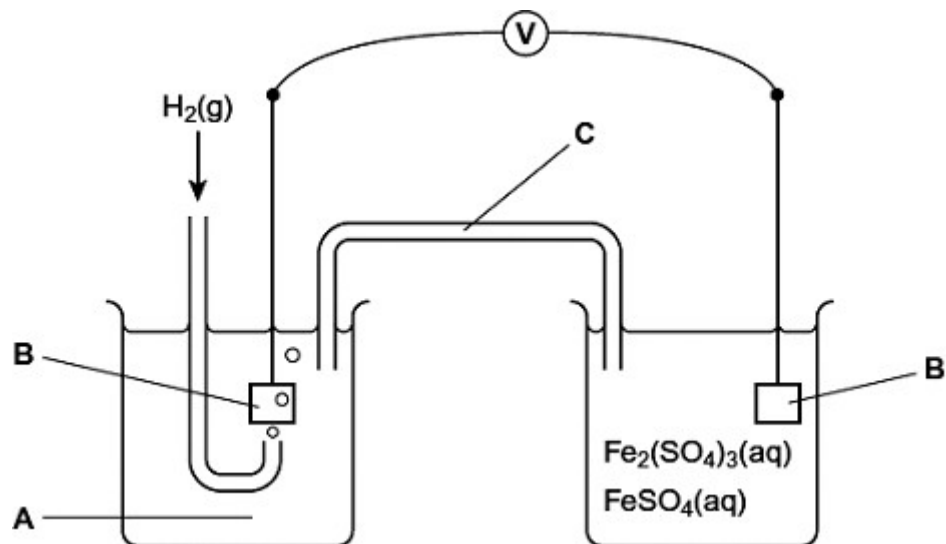
- (d) Potassium manganate(VII) is an oxidising agent.
Suggest **one** reason why a $0.0200 \text{ mol dm}^{-3}$ solution of potassium manganate(VII) does **not** need to be kept away from flammable material.

(1)

(Total 9 marks)

Q19.

The diagram below shows a cell that can be used to measure the standard electrode potential for the half-reaction $\text{Fe}^{3+}(\text{aq}) + \text{e}^- \longrightarrow \text{Fe}^{2+}(\text{aq})$. 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.

(1)

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

(1)
(Total 12 marks)

Q20.

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

Electrode half-equation	$E^\ominus / \text{V}$
$\text{Cl}_2(\text{aq}) + 2\text{e}^- \longrightarrow 2\text{Cl}^-(\text{aq})$	+1.36
$2\text{HOCl}(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{Cl}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$	+1.64
$\text{H}_2\text{O}_2(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow 2\text{H}_2\text{O}(\text{l})$	+1.77
$\text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{H}_2\text{O}_2(\text{aq})$	+0.68
$\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \longrightarrow 2\text{H}_2\text{O}(\text{l})$	+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 _____

Equation _____

(2)

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

(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.

(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 _____

Equation _____

(2)

(Total 8 marks)

Q21.

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	$E^\ominus / \text{V}$
$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn}(\text{s})$	-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.

_____ (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	$E^\ominus / \text{V}$
$\text{PbO}_2(\text{s}) + 3\text{H}^+(\text{aq}) + \text{HSO}_4^-(\text{aq}) + 2\text{e}^- \longrightarrow \text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l})$	+1.69
$\text{PbSO}_4(\text{s}) + \text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{Pb}(\text{s}) + \text{HSO}_4^-(\text{aq})$	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.

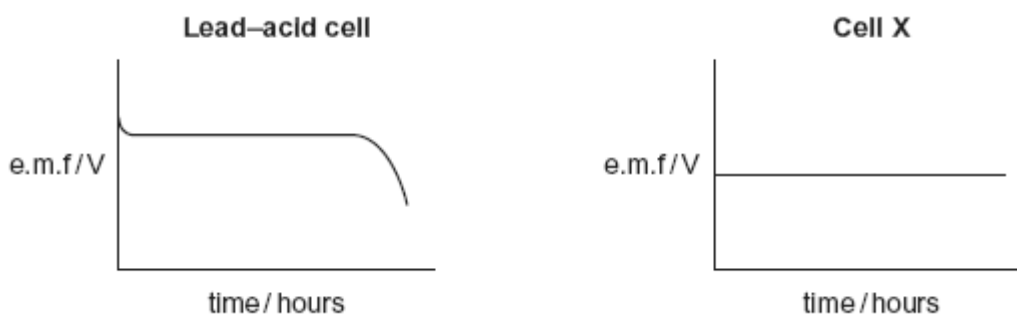
Use this information to calculate the missing electrode potential for the half-equation shown in the table.

(1)

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

(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**.

(2)

(Total 17 marks)

Q22.

The table below shows some standard electrode potentials.

			$E^\ominus / \text{V}$
$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^-$	$\longrightarrow$	$\text{Mn}^{2+} + 4\text{H}_2\text{O}$	+1.51
$\text{Cl}_2(\text{g}) + 2\text{e}^-$	$\longrightarrow$	$2\text{Cl}^-(\text{aq})$	+1.36
$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^-$	$\longrightarrow$	$2\text{Cr}^{3+} + 7\text{H}_2\text{O}$	+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.

(Total 2 marks)

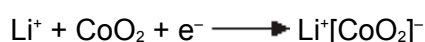
Q23.

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



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.



- (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 _____

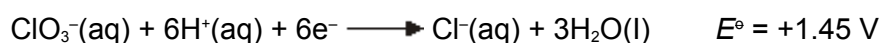
Property 2 _____

(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.



- (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 _____

Q24.

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°/V
$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$	+1.33
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq})$	+0.77
$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	0.00
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Fe}(\text{s})$	-0.44
$\text{Li}^+(\text{aq}) + \text{e}^- \rightarrow \text{Li}(\text{s})$	-3.04

- (a) Describe a standard hydrogen electrode.

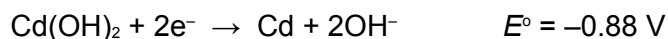
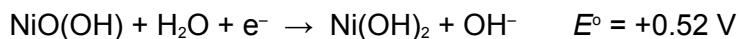
(4)

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



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.



- (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)

Q26.

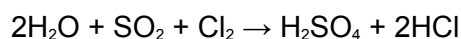
- (a) The term oxidation was used originally to describe a reaction in which a substance gained oxygen. The oxygen was provided by the oxidising agent. Later the definition of oxidation was revised when the importance of electron transfer was recognised.

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

Reaction 1 with HgO



Reaction 2 with chlorine



- (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 table below to answer the questions which follow.

	<i>E</i> / 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

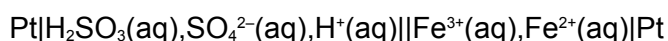
- (i) An excess of potassium manganate(VII) was added to a solution containing $V^{2+}(aq)$ ions. Determine the vanadium species present in the solution at the end of this reaction. State the oxidation state of vanadium in this species and write a half-equation for its formation from $V^{2+}(aq)$.

Vanadium species present at the end of the reaction _____

Oxidation state of vanadium in the final species _____

Half-equation _____

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



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.

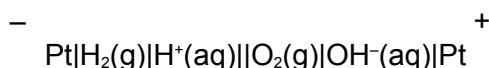
Cell e.m.f. _____

Change in cell e.m.f. _____

Explanation _____

(7)

- (c) Consider the cell below



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

-
-
-
- (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)

Q27.

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

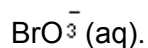
Standard electrode potentials	$E^\ominus / \text{V}$
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \longrightarrow \text{Fe}^{2+}(\text{aq})$	+0.77
$\text{Cl}_2(\text{g}) + 2\text{e}^- \longrightarrow 2\text{Cl}^-(\text{aq})$	+1.36
$2\text{BrO}_3^-(\text{aq}) + 12\text{H}^+(\text{aq}) + 10\text{e}^- \longrightarrow \text{Br}_2(\text{aq}) + 6\text{H}_2\text{O}(\text{l})$	+1.52
$\text{O}_3(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{O}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$	+2.08
$\text{F}_2\text{O}(\text{g}) + 2\text{H}^+(\text{aq}) + 4\text{e}^- \longrightarrow 2\text{F}^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$	+2.15

Each of the above can be reversed under suitable conditions.

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

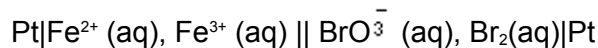
- (ii) Identify the most powerful oxidising agent in the table.

- (iii) Identify **all** the species in the table which can be oxidised in acidic solution by



(4)

(b) The cell represented below was set up.



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

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

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

Change in concentration _____

Explanation _____

(6)

(Total 10 marks)

Q28.

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

					$E^\ominus/\text{V}$
$\text{Zn}^{2+}(\text{aq})$	+	2e^-	$\rightarrow$	$\text{Zn}(\text{s})$	-0.76
$\text{V}^{3+}(\text{aq})$	+	e^-	$\rightarrow$	$\text{V}^{2+}(\text{aq})$	-0.26
$\text{SO}_4^{2-}(\text{aq}) + 2\text{H}^+(\text{aq})$	+	2e^-	$\rightarrow$	$\text{SO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l})$	+0.17
$\text{VO}^{2+}(\text{aq}) + 2\text{H}^+(\text{aq})$	+	e^-	$\rightarrow$	$\text{V}^{3+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$	+0.34
$\text{Fe}^{3+}(\text{aq})$	+	e^-	$\rightarrow$	$\text{Fe}^{2+}(\text{aq})$	+0.77
$\text{VO}_2^+(\text{aq}) + 2\text{H}^+(\text{aq})$	+	e^-	$\rightarrow$	$\text{VO}^{2+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$	+1.00
$\text{Cl}_2(\text{aq})$	+	2e^-	$\rightarrow$	$2\text{Cl}^-(\text{aq})$	+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 $\text{VO}_2^+(\text{aq})$ to $\text{VO}^{2+}(\text{aq})$ but will **not** reduce $\text{VO}^{2+}(\text{aq})$ to $\text{V}^{3+}(\text{aq})$,

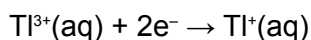
- (ii) a species which, in acidic solution, will oxidise $\text{VO}^{2+}(\text{aq})$ to $\text{VO}_2^+(\text{aq})$.

(2)

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

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

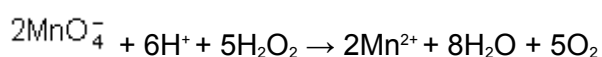
- (i) Deduce the standard electrode potential for the following half-reaction.



- (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 $0.0200 \text{ mol dm}^{-3}$ solution of potassium manganate(VII). The overall equation for the reaction is given below.



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

- (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.

(8)
(Total 14 marks)

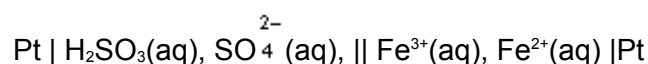
Q29.

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
$IO_3^-(aq) + 12H^+(aq) + 10e^- \rightarrow I_2(aq) + 6H_2O(l)$	+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.

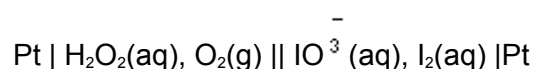


(i) Calculate the e.m.f. of this cell.

(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.



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

(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 $\text{V}^{2+}(\text{aq})$ 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 $\text{V}^{2+}(\text{aq})$.

Vanadium species present at end of reaction _____

Oxidation state of vanadium in final species _____

Half-equation _____

(3)

(Total 12 marks)

Q30.

In this question consider the data below.

	$E^\ominus / \text{V}$
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$	+0.80
$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	0.00
$\text{Pb}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Pb}(\text{s})$	-0.13

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

- A** 0.93 V
B 0.67 V
C -0.67 V

D -0.93 V

(Total 1 mark)

Q31.

In this question consider the data below.

	$E^\ominus / \text{V}$
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$	+0.80
$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	0.00
$\text{Pb}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Pb}(\text{s})$	-0.13

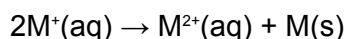
The e.m.f. of the cell $\text{Pt}(\text{s}) \mid \text{H}_2(\text{g}) \mid \text{H}^+(\text{aq}) \parallel \text{Ag}^+(\text{aq}) \mid \text{Ag}(\text{s})$ would be increased by

- A increasing the concentration of $\text{H}^+(\text{aq})$.
- B increasing the surface area of the Pt electrode.
- C increasing the concentration of $\text{Ag}^+(\text{aq})$.
- D decreasing the pressure of $\text{H}_2(\text{g})$.

(Total 1 mark)

Q32.

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



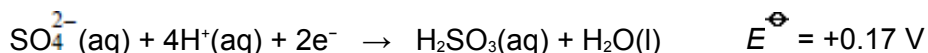
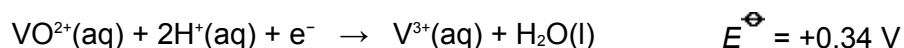
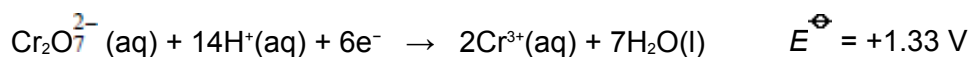
The table below contains $E^\ominus$ data for copper and mercury species.

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

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

- A Both $\text{Cu}(\text{l})$ and $\text{Hg}(\text{l})$ undergo disproportionation.
- B Only $\text{Cu}(\text{l})$ undergoes disproportionation.
- C Only $\text{Hg}(\text{l})$ undergoes disproportionation.
- D Neither $\text{Cu}(\text{l})$ nor $\text{Hg}(\text{l})$ undergoes disproportionation.

(Total 1 mark)

Q33.

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

- A** 40 cm³ of a 0.10 mol dm⁻³ solution of Cr₂O₇²⁻(aq)
- B** 80 cm³ of a 0.30 mol dm⁻³ solution of Fe³⁺(aq)
- C** 50 cm³ of a 0.24 mol dm⁻³ solution of V³⁺(aq)
- D** 50 cm³ of a 0.24 mol dm⁻³ solution of H₂SO₃(aq)

(Total 1 mark)

Q34.

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

Standard electrode potentials	$E^\ominus/\text{V}$
$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	0.00
$\text{Br}_2(\text{aq}) + 2\text{e}^- \rightarrow 2\text{Br}^-(\text{aq})$	+1.09
$2\text{BrO}_3^-(\text{aq}) + 12\text{H}^+(\text{aq}) + 10\text{e}^- \rightarrow \text{Br}_2(\text{aq}) + 6\text{H}_2\text{O}(\text{l})$	+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 _____

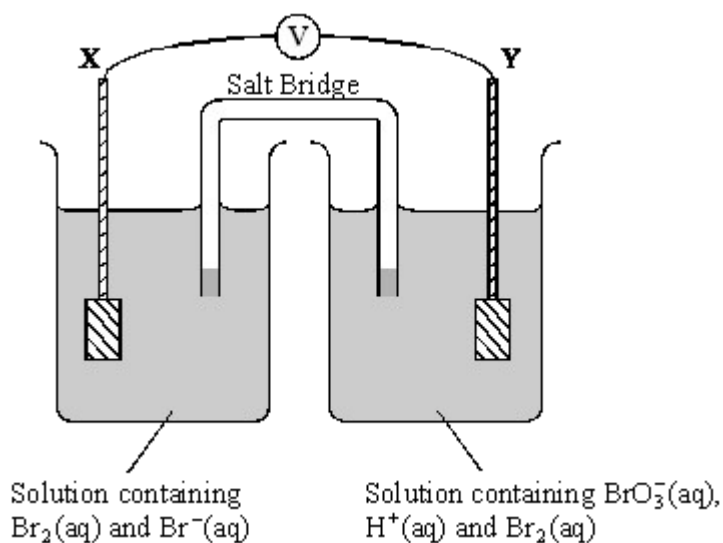
Hydrogen gas pressure _____

(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 _____

- (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 _____

Q35.

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

	$E^\ominus / \text{V}$
$\text{MnO}_4^- (\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+ 1.52
$\text{Cr}_2\text{O}_7^{2-} (\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$	+ 1.33
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq})$	+ 0.77
$\text{Cr}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Cr}^{2+}(\text{aq})$	- 0.41
$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn}(\text{s})$	- 0.76

The most powerful oxidising agent in the table is

- A** $\text{Mn}^{2+}(\text{aq})$
- B** $\text{Zn}(\text{s})$
- C** $\text{MnO}_4^{-}(\text{aq})$
- D** $\text{Zn}^{2+}(\text{aq})$

(Total 1 mark)

Mark schemes

Q1.

- (a) It has mobile ions / ions can move through it / free ions

Do not allow movement of electrons.

1

- (b) (+) 0.18 V

1

- (c) The concentration is not 1.0 (mol dm^{-3})

1

- (d) $\text{Cu (s)} \mid \text{Cu}^{2+}(\text{aq}) \parallel \text{Cu}^{2+}(\text{aq}) \mid \text{Cu(s)}$

1

- (e) (Concentration) increases or ($[\text{Cu}^{2+}]$ ions) increase

Mark independently

1

The $[\text{Cu}^{2+}]$ ions in the two solutions become equal/same

Not, concentrations are constant

1

[6]

Q2.

- (a) Fe^{2+}

Accept any Fe(II) compound – correct formula or name

1

$$E^\ominus \text{VO}_2^+ / \text{VO}^{2+} > E^\ominus \text{Fe}^{3+} / \text{Fe}^{2+} > E^\ominus \text{VO}^{2+} / \text{V}^{3+}$$

If calculations of EMF are provided producing EMFs = 0.23(V) and -0.43(V), with a comment, allow M2

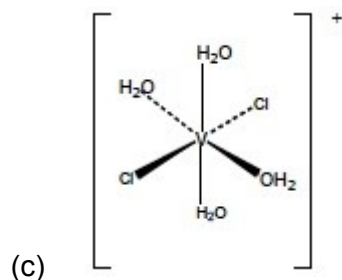
allow $E^\ominus \text{Fe}^{3+} / \text{Fe}^{2+}$ value of +0.77 is between the $E^\ominus$ values for the electrode half-equations containing the V species or wtte

1

- (b) (+) 4

IV or four

1



Ignore absence of charge

Wedges, dotted lines and [] not required

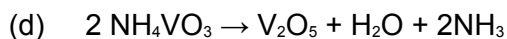
Do not penalise bond from H to V (in water ligands)

1

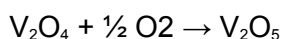
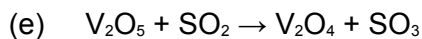
Cis/trans

allow E/Z, geometric and stereo(isomerism)

1

*Accept multiples**Ignore state symbols*

1

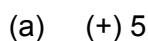
*Both equations needed for 1 mark in this order**Allow multiples*

1

[7]

Q3.**B**

[1]

Q4.

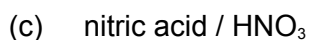
1



1

Allow Roman numerals*Allow 2Cl^-* *Do not allow chlorine / Cl / Cl_2* *Ignore (aq)*

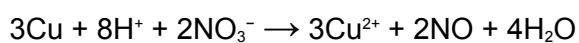
1

*If not nitric acid then CE = 0**If NO_3^- ions identified, lose M1 and mark on**Allow $0.96\text{V} > 0.34\text{V}$*

1

the $E^\ominus \text{NO}_3^-$ (/NO) $>$ $E^\ominus \text{Cu}^{2+}$ (/Cu) or in words*Allow NO_3^- is a better oxidising agent than Cu^{2+}* *Allow NO_3^- has a more positive $E^\ominus$ than Cu^{2+}*

1

*Allow $3\text{Cu} + 8\text{HNO}_3 \rightarrow 3\text{Cu}(\text{NO}_3)_2 + 2\text{NO} + 4\text{H}_2\text{O}$*

1

EMF for the reaction is 0.62(V)

1

[7]

Q5.

(a) Cl_2

1

(b) $\text{Pt}|\text{H}_2|\text{H}^+||\text{Sn}^{2+}|\text{Sn}$

Award 1 mark for species and 1 mark for correct order including Pt

Ignore state symbols

2

(c) 0.94 V

1



1

(d) $\text{Ag}^+ + \text{Fe}^{2+} \rightarrow \text{Ag} + \text{Fe}^{3+}$

1

[6]

Q6.

(a) The ions in the ionic substance in the salt bridge move through the salt bridge

1

To maintain charge balance / complete the circuit

1

(b) F^-

1

(c) $E^\ominus \text{SO}_4^{2-} / \text{SO}_2 < E^\ominus \text{Br}_2 / \text{Br}^-$

Allow correct answer expressed in words, eg electrode potential for sulfate ions / sulfur dioxide is less than that for bromine / bromide

1

(d) 1.23 (V)

1

(e) A fuel cell converts more of the available energy from combustion of hydrogen into kinetic energy of the car / an internal combustion engine wastes more (heat) energy

1

[6]

Q7.

A

Q8.

- (a) Magnesium

If Mg^{2+} then $CE = 0/2$

1

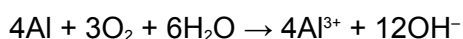
It is the most powerful reducing agent / it has the most negative $E^\ominus$ / forms the cell with iron with the biggest emf

1

- (b)
- $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

*Allow fractions and multiples.**Ignore state symbols.*

1

*Allow fractions and multiples.**Ignore state symbols.*

1

$$E = 0.40 - (-1.66) = \underline{2.06} \text{ (V)}$$

Lose this mark if answer is -2.06

1

- (c) Non-standard conditions (T, p or concentration)

Current flow affects voltage

 Al^{3+} ions may react in alkaline conditions / with OH^- / form $Al(OH)_3$ *Max 2 from 3*

2

- (d) The iron foil does not lose mass as it is protected by the operation of the sacrificial anode

1

- (e) A coating developed on the aluminium preventing further reaction

1

[9]

Q9.

- (a) (i)
- Zn^{2+}

 $Zn^{2+}(aq)$ *Apply List*

1

- (ii) 298 K / 25°C

*Ignore pressure**Ignore standard conditions**Ignore state symbols*

1

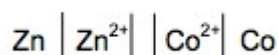
(Solutions at) unit concentration / 1 mol dm⁻³ (of Zn²⁺)

Ignore references to S.H.E

1

- (b) Identifying it is the Zn/Zn²⁺ and Co²⁺/Co half cells

1



*Correct order with phase boundaries and salt bridge correct,
no Pt*

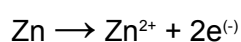
If this is correct it scores M1 and M2

Allow double dashed line for salt bridge

Extra phase boundaries loses M2

Ignore state symbols

1



M3 independent

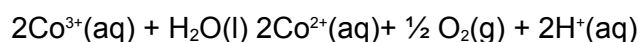
Allow -2e^{-} on LHS

1

- (c) Co³⁺

Mark independently

1



Ignore state symbols

Allow multiples

1

Oxygen /O₂

Allow $\frac{1}{2}\text{O}_2$

1

- (d) ***E***[°] (O₂ | H₂O) electrode ***E***[°] (O₂ | H₂O)
OR the ***E***[°] (Au⁺ | Au) electrode potential is more positive than the ***E***[°] (O₂ | H₂O) electrode
OR The emf (for the reaction of Au and oxygen) is -0.45 V (and therefore not spontaneous)

Mark independently

1

So oxygen is unable to oxidise gold

Ignore references to water

Allow gold cannot reduce oxygen

1

[11]

Q10.

- (a) Electron acceptor / gains electrons
do not allow electron pair acceptor

(b) Fe^{2+} ions

Fe^{2+} / Fe or Fe^{2+} or it has smallest / most negative electrode potential / E°

Do not allow Fe / Fe^{2+}

Cannot score M2 if M1 incorrect

(c) $\text{Pt}|\text{H}_2|\text{H}^+||\text{Ag}^+|\text{Ag}$

M1 for H_2 H^+ Ag^+ Ag in correct order

allow dashed phase boundaries

2H^+ loses one mark (M2)

M2 for Pt correct and correct phase boundaries

Ignore state symbols. M1 must be correct to score M2

If answer correct but all in reverse order allow 1 mark out of two

Any **two** correct conditions

- 298 K / 25 °C
- 100 kPa
- both solutions of unit concentration
- zero current

Allow 1 bar

Do not apply list principle, mark correct answers.

(d) $E_{\text{Au}^+ / \text{Au}} > E_{\text{O}_2 / \text{H}_2\text{O}}$ OR e.m.f. / $E_{\text{cell}} = 0.45 \text{ V}$

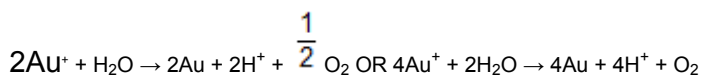
*If both species in electrode given, must be in correct order
i.e. Au^+ / Au*

Au^+ (ions) oxidise water OR water reduces Au^+ (ions)

Allow water donates electrons to Au^+

Gold metal / solid / precipitate **OR** bubbles / effervescence of (oxygen gas) / gas produced

Penalise incorrect observations



Allow multiples

(e) (i) 1.24 (V)

Do not allow -1.24

(ii) Chloride ions / Cl^- react with / form a precipitate with silver ions / Ag^+ /

form AgCl

Penalise reaction of chloride ions with iron ions or iron

1

- (f) $E_{O_2 (/ H_2O)} > E_{Fe^{3+} (/ Fe^{2+})}$ (or e.m.f / $E_{cell} = 0.46 \text{ V}$)

Species in electrode if all given must be in correct order

1

Therefore the iron(II) ions are oxidised (or converted) into iron(III) ions (by oxygen)

If chloride ions oxidised to chlorine, lose M2

M2 can be obtained or lost from equation.

Ignore observations.

1

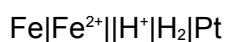
[15]

Q11.

- (a) $Pt|H_2|H^+||Fe^{2+}|Fe$

Allow 1 for correct order of symbols but lose second mark for a wrong phase boundary(s) / Pt missing / extra Pt on RHS, additional phase boundary

Note, allow one mark only for correct symbol in reverse:



Allow dashed lines for salt bridge

Ignore state symbols

Ignore 2 if used before H^+

2

- (b) Electron donor

Allow (species that) loses electrons

Do not allow reference to electron pairs

1

- (c) Cl_2 / chlorine

If M1 blank or incorrect cannot score M2

1

(Species on RHS / electron donor) has most positive / largest $E^\ominus$ / has highest potential

Do not allow reference to e.m.f. or $E_{(cell)}$

1

- (d) (i) Cl / chlorine

1

- (ii) Chlorine +1 to chlorine 0

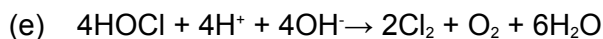
CE if chlorine not identified in part (i)

Allow chlorine +1 to chlorine -1 (in Cl^-)

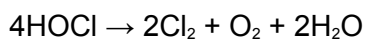
Allow oxidation state decreases by one OR two

Allow oxidation state changes by -1 OR -2

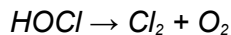
1



OR



Allow one mark for any incorrect equation that shows



Allow multiples

Ignore state symbols

Penalise one mark for uncanceled or uncombined species

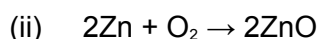
(eg $\text{H}_2\text{O} + \text{H}_2\text{O}$ instead of $2\text{H}_2\text{O}$)

2

(f) (i) e.m.f. = $0.40 - (-1.25) = \underline{1.65}$ (V) / $\underline{+1.65}$ (V)

Allow -1.65 (V)

1



Allow multiples

Ignore state symbols

Do not allow uncanceled species

If more than one equation given, choose the best

1

(iii) **A** / stainless lid

If M1 incorrect or blank CE=0

1

O₂ (electrode) has a more positive $E^\ominus$ / oxygen (electrode) requires / gains electrons from external circuit

Or reference to the overall equation and a link to electrons going into A

Allow oxygen is reduced and reduction occurs at the positive electrode

OR Zinc (electrode) has more negative $E^\ominus$

Do not allow reference to e.m.f. or $E(\text{cell})$

1

(iv) (Cell) reaction(s) cannot be reversed / zinc oxide cannot be reduced to zinc by passing a current through it / zinc cannot be regenerated

Allow danger from production of gas / oxygen produced / hydrogen produced

1

[14]

Q12.

(a) Platinum electrode

1

Solution in beaker is a mixture of named soluble iron(II) compound and named soluble iron(III) compound

Allow correct formulae for the iron compounds.

Concentrations of Fe(II) and Fe(III) ions are both 1 mol dm⁻³

Ignore any references to temperature.

If eg Fe₂(SO₄)₃ used then concentration must be 0.5

1

- (b) Purpose: Allow movement of ions between electrodes

Allow to maintain an electric circuit.

Do not allow reference to movement of electrons in salt bridge.

1

Requirement: Must not react with the electrolyte / ions in solution

Do not allow 'must not react' without further qualification.

1

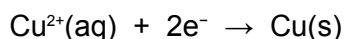
[5]

Q13.

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

*If equations reversed, allow **M1** only.*

1



Ignore state symbols.

1

- (b) Moles of copper(II) reacted = $(100 / 1000) \times 0.5 = 0.05$

1

Moles of zinc reacted = 0.05

1

Mass of zinc lost = $0.05 \times 65.4 = 3.27 \text{ g}$

*Correct final answer without working scores **M3** only.*

1

- (c) Allow cell to discharge until [Cu²⁺] is 0.5

Alternative: Allow cell to discharge completely.

1

Confirmed by colorimetric measurement or other suitable method

Solution colourless or use of chemical test to determine absence of copper(II)

1

Weigh the Zn electrode before and after the experiment

Weigh Zn electrodes before and after and halve the mass change.

1

[8]

Q14.

- (a) It has mobile ions / ions can move through it / free ions

Do not allow movement of electrons.

Allow specific ions provided they are moving but do not react.

1

- (b) Chloride ions react with copper ions / Cu^{2+} **OR** $[\text{CuCl}_4]^{2-}$ formed
If incorrect chemistry, mark = 0

1

- (c) The Cu^{2+} ions / CuSO_4 in the left-hand electrode more concentrated
Allow converse.

1

So the reaction of Cu^{2+} with 2e^- will occur (in preference at) left-hand electrode
/ $\text{Cu} \rightarrow \text{Cu}^{2+} + \text{electrons}$ at right-hand electrode

Allow left-hand electrode positive / right-hand electrode negative.

Also reduction at left-hand electrode / oxidation at right-hand electrode.

Also left-hand electrode has oxidising agent / right-hand electrode has reducing agent.

Allow E left-hand side $>$ E right-hand side

1

- (d) (Eventually) the copper ions / CuSO_4 in each electrode will be at the same concentration

1

- (e) (i) -3.05 (V)
Must have minus sign.
 -3.05 only.

1

- (ii) $\text{LiMnO}_2 \rightarrow \text{Li} + \text{MnO}_2$ correct equation
Allow 1 for reverse equation.
Allow multiples.

1

Correct direction

If Li^+ not cancelled but otherwise correct, max = 1

If electrons not cancelled, CE = 0

$\text{LiMnO}_2 \rightarrow \text{Li} + \text{MnO}_2$ scores 2

$\text{Li}^+ + \text{LiMnO}_2 \rightarrow \text{Li}^+ + \text{Li} + \text{MnO}_2$ scores 1

$\text{Li} + \text{MnO}_2 \rightarrow \text{LiMnO}_2$ scores 1

1

- (iii) Electricity for recharging the cell may come from power stations burning (fossil) fuel

Allow any reference to burning (of carbon-containing) fuels.

Note combustion = burning.

1

[9]

Q15.

- (a) To remove the oxide layer on the aluminium
Do not allow 'cleaning' or 'removal of grease'.
Do not allow 'removal of impurities' without qualification. 1

- (b) An appropriate method for delivering H₂ gas over a Pt electrode
Need H₂ gas and Pt electrode labelled (allow gas delivered directly below the electrode). 1

The Pt electrode must clearly be in contact with a solution of a named acid.
Ignore any concentration or pressure values.
Ignore absence of bubbles.
Allow if electrode is below outer acid level. 1

- (c) The carbonate ion reacts with the acid (in the SHE) / reaction between carbonate and Al³⁺
Lose this mark if aluminium carbonate formed but mark on. 1

Reaction given (either equation or products specified)

OR H⁺ / Al³⁺ concentrations change / cell e.m.f. altered 1

[5]

Q16.

- (a) Diagram of an Fe³⁺ / Fe²⁺ electrode that includes the following parts labelled:
 Solution containing Fe²⁺ and Fe³⁺ ions 1

Platinum electrode connected to one terminal of a voltmeter
Must be in the solution of iron ions (one type will suffice) 1

Salt bridge
Do not allow incorrect material for salt bridge and salt bridge must be in the solution (ie it must be shown crossing a meniscus) 1

298 K and 100 kPa / 1 bar 1

all solutions unit / 1 mol dm⁻³ concentration
Allow zero current / high resistance voltmeter as alternative to M4 or M5
Ignore hydrogen electrode even if incorrect 1

- (b) Cu²⁺ + Fe → Cu + Fe²⁺
Ignore state symbols 1

Fe|Fe²⁺||Cu²⁺|Cu correct order
Allow Cu|Cu²⁺||Fe²⁺|Fe

1

Phase boundaries and salt bridge correct, no Pt
Allow single / double dashed line for salt bridge
Penalise phase boundary at either electrode end
Can only score M3 if M2 correct

1

Copper electrode
Allow any reference to copper

1

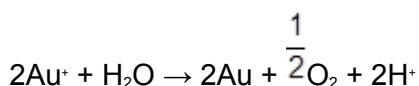
- (c) $E^\ominus \text{Au}^+ / \text{Au} > E^\ominus \text{O}_2 / \text{H}_2\text{O}$
Allow $E_{\text{cell}} / \text{e.m.f.} = 0.45 \text{ V}$
Allow $1.68 > 1.23$

1

So Au⁺ ions will oxidise water / water reduces Au⁺

QoL

1



Allow multiples

1

- (d) $E^\ominus \text{Ag}^+ / \text{Ag} > E^\ominus \text{Fe}^{2+} / \text{Fe}$
Allow $E_{\text{cell}} / \text{e.m.f.} = 1.24$
Allow $0.80 > -0.44$

1

And $E^\ominus \text{Ag}^+ / \text{Ag} > E^\ominus \text{Fe}^{3+} / \text{Fe}^{2+}$
Allow $E_{\text{cell}} / \text{e.m.f.} = 0.03$
Allow $0.80 > 0.77$

1

So silver ions will oxidise iron (to iron(II) ions) and then oxidise Fe(II) ions (further to Fe(III) ions producing silver metal)

Allow Ag⁺ ions will oxidise iron to iron(III)

1

[15]

Q17.

- (a) Electron acceptor / gains electrons / takes electrons away
Do not allow electron pair acceptor / gain of electrons / definition of redox (QWC)

1

- (b) Cd(OH)₂
Do not allow 'Cd(OH)₂/Cd'

Species (on LHS) with the least positive/most negative electrode potential / lowest E / smallest E

Only allow this mark if M1 answer given correctly or blank

Do not allow negative emf

1

(c) (i) 1.5 (V) / 1.50

1

(ii) $2\text{MnO}_2 + 2\text{H}_2\text{O} + \text{Zn} \rightarrow 2\text{MnO}(\text{OH}) + 2\text{OH}^- + \text{Zn}^{2+}$

Ignore state symbols

e^- must be cancelled

(take care that Zn^{2+} is on RHS)

1

(iii) Allows ions to pass (through it) or words to that effect

Penalise passage of electrons

Allow mention of particular ions

1

(iv) Allows electrons to flow / makes electrical contact / conductor

Allow acts as an (inert) electrode / anode / cathode

1

(v) Zn is 'used up' / has reacted / oxidised

Allow idea that zinc reacts

Do not allow just zinc corrodes

1

(d) (i) 3 / +3 / III

1

$2\text{Ni}(\text{OH})_2 + \text{Cd}(\text{OH})_2 \rightarrow 2\text{NiO}(\text{OH}) + \text{Cd} + 2\text{H}_2\text{O}$

For correct nickel and cadmium species in correct order

(allow H_2O missing and OH^- not cancelled)

1

For balanced equation (also scores M2)

Allow max 1 for M2 and M3 if correct balanced equation but reversed.

Ignore state symbols

1

(ii) Metal / metal compounds are re-used / supplies are not depleted / It (the cell) can be re-used

Allow does not leak / no landfill problems / less mining / less energy to extract metals / less waste

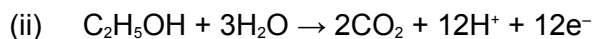
Do not allow less CO_2 unless explained

1

(e) (i) $\text{C}_2\text{H}_5\text{OH} + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O}$

Allow $\text{C}_2\text{H}_6\text{O}$

1



Allow C₂H₆O

1

(iii) (+)0.23 (V)

1

(iv) CO₂ released by combustion / fermentation / fuel cell / reaction with water
Can be answered with the aid of equations

1

(atmospheric) CO₂ taken up in photosynthesis

1

[17]

Q18.

(a) Manganate would oxidise / react with Cl⁻

1

Because $E^\ominus$ for MnO_4^- is more positive than that for Cl^2 / $1.51 - 1.36 = +0.15$ (V)

Must refer to data from the table for M2.

1

(b) Moles of H^+ = $25 \times 0.0200 \times 8 / 1000 = 4.00 \times 10^{-3}$

1

Moles of H_2SO_4 = 2.00×10^{-3} ($4.00 \times 10^{-3} / 2$)

Allow consequential marking on incorrect moles of H^+

1

Volume H_2SO_4 = $4.00 \text{ (cm}^3\text{)}$ ($2.00 \times 10^{-3} \times 1000 / 0.500$)

Allow consequential marking on incorrect moles of H_2SO_4

Accept 4 cm^3 .

8 cm^3 scores 2 marks.

Do not penalise precision.

Correct answer without working scores M3 only.

1

(c) (i) $\text{MnO}_4^- + 4\text{H}^+ + 3\text{e}^- \rightarrow \text{MnO}_2 + 2\text{H}_2\text{O}$

Allow multiples, including fractions.

Ignore state symbols.

1

(ii) Can't see end point due to brown colour

1

Larger titre (than expected)

Allow the idea that with two reactions can't make use of titre in calculations.

Do not allow 'an inaccurate result' without qualification.

1

(d) Solution (very) dilute / lots of water

1

[9]

Q19.

(a) HCl 1.0 mol dm⁻³

Allow H₂SO₄ 0.5 mol dm⁻³

Allow HNO₃ 1.0 mol dm⁻³

Allow name or formula

Concentration can be given after "conditions"

1

(Hydrogen at) 100kPa / 1 bar

1

298 K

1

(b) Pt / Platinum

Mark on if no answer for M1

If wrong answer for M1, only mark on if electrode is Au, Ag, Pb or Ti

1

Inert / unreactive / does not create a potential difference

1

Conducts electricity / allows electron flow / conducts / conductor

1

(c) KCl

Allow NaCl, KNO₃, Na₂SO₄ etc NOT NH₄Cl

1

Does not react with either electrode / solution in electrode

Allow unreactive / inert

1

Ions can move

Allow conducts electricity / electrical connection / carries charge

Do not allow just connects / completes the circuit

Do not allow conducts / carries electrons

Mark these independently

1

(d) Pt|H₂|H⁺||Fe³⁺,Fe²⁺|Pt

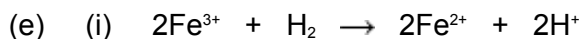
Ignore state symbols

Order must be correct

| must be correct but allow | instead of , separating Fe³⁺ from Fe²⁺

Allow , instead of | separating H₂ and H⁺

1



Allow multiples

1

(ii) The Fe^{3+} ions would be used up / reaction completed

Answer must relate to reactants in (e)(i) equation if given

Allow reactant / reactants used up

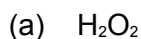
Do not allow concentration of Fe^{3+} decreases

Allow concentration of Fe^{3+} falls to zero

1

[12]

Q20.



Ignore state symbols

1



Allow potential for chlorine/ Cl_2 greater than for oxygen/ O_2

Allow $1.36 > 1.23$ / $E_{\text{cell}} = 0.13$

1



Allow multiples

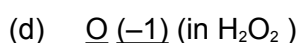
Allow + HCl

1

(c) Activation energy is high / light/UV provides the activation energy / light breaks chlorine molecule / Cl–Cl bond

If light used to break Cl–Cl bond award 1 mark and ignore product e.g. Cl^-

1



Must give oxidation state of O in $\text{H}_2\text{O}_2 = -1$

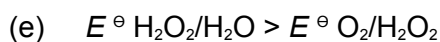
1

Changes to O(–2) (in water)

Must give oxidation state of O in water = –2

CE = 0/2 if refers to oxidation state of H changing

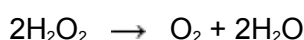
1



Allow stated in words

Allow $1.77 > 0.68$ / $E_{\text{cell}} = 1.09$

1



Allow multiples

H^+ and e^- must be cancelled

1

[8]

Q21.

- (a) loses electrons / donates electrons

penalise donates electron pair

1

- (b) Zn

1

(most) negative E° / lowest E° / least positive*can only score M2 if M1 correct**do not allow e.m.f instead of E°*

1

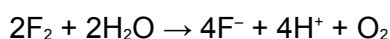
- (c)
- $E^\circ \text{F}_2 / \text{F}^- > E^\circ \text{O}_2 / \text{H}_2\text{O}$

or e.m.f is positive or e.m.f = 1.64 V

1

Fluorine reacts to form oxygen (can score from equation in M3 even if equation unbalanced provided no contradiction)
 or fluorine oxidises water
 or fluorine is a more powerful oxidising agent than oxygen

1

*allow 4HF in equation**balanced equation scores M2 and M3*

1

- (d) (i) order correct Zn Zn
- ²⁺
- Ag
- ₂
- O Ag or reverse of this order

ignore ss, H⁺ and H₂O, no. of moles

1

all phase boundaries correct

*allow Zn|Zn²⁺||Ag₂O,Ag**or Zn|Zn²⁺||Ag₂O|H⁺|Ag for M1 & M2*e.g. Zn|Zn²⁺||Ag₂O|Ag or Ag|Ag₂O||Zn²⁺|Zn scores 2*M2 cannot be gained unless M1 scored**allow H⁺ either side of Ag₂O with comma or |**for M2 penalise*

- *wrong phase boundary (allow dashed lines for salt bridge)*
- *Pt*
- *use of + (from half equation)*
- *water/H⁺ outside Ag in Ag electrode*

1

- (ii) 1.1 (V)

*Allow no units, penalise wrong units**allow correct answer even if no answer to (d)(i) or answer to (d)(i) incorrect**allow -1.1 if silver electrode on Left in (d)(i) even if the species are in the wrong order.*

1

- (iii) Reaction(s) not reversible or H₂O electrolyses
do not allow hard to reverse
mention of primary cell is not enough to show that reaction(s) are irreversible 1
- (e) (i) -0.46 (V)
Allow no units, penalise wrong units 1
- (ii) $2\text{PbSO}_4 + 2\text{H}_2\text{O} \rightarrow \text{Pb} + \text{PbO}_2 + 2\text{HSO}_4^- + 2\text{H}^+$
 lead species correct on correct sides of equation 1
 equation balanced and includes H₂O,
 HSO₄⁻ and H⁺ (or H₂SO₄)
allow ions / species must be fully cancelled out or combined
allow 1/2 for balanced reverse equation 1
- (f) (i) reagents / PbO₂ / H₂SO₄ / acid / ions used up
 (or concentration decreases) 1
- (ii) fuel cell
Ignore any other words 1
- (iii) reagents / fuel supplied continuously 1
 concentrations (of reagents) remain constant 1

[17]

Q22.

MnO₄⁻ will oxidise the chloride ion / reaction of MnO₄⁻ and Cl⁻ feasible
Accept converse argument with Cr₂O₇²⁻
Accept calculations of overall E° values.

1

Larger volume needed

1

[2]

Q23.

- (a) (i) Co/Cobalt
If Co or Cobalt not given CE = 0
ignore case in symbol for Co

1

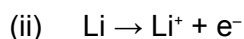
(+) 4

1

(+) 3

Allow 4 and 3 in either order

1



Ignore state symbols

Allow e without -ve sign

Do not allow equilibrium sign

1

(iii) Platinum is a conductor

1

(Platinum is) unreactive/inert

Ignore mention of surface area or catalyst

Allow 2 marks if two properties given on one answer line

Apply list principle to contradictions/wrong answers

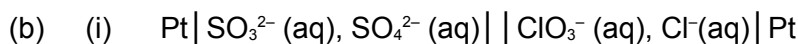
Do not allow platinum resists corrosion

1

(iv) Li reacts with water/forms lithium hydroxide

Allow water breaks down (or is electrolysed) on re-charge

1



State symbols as ', ' not necessary

Allow | in place of ', ' NOT ', ' in place of |

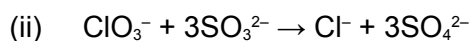
Ignore H^+ and H_2O

Deduct one mark for each mistake (e.g. Pt missed twice counts as two mistakes)

Allow reverse order for whole cell



2



1

Oxidising agent ClO_3^-

1

Reducing agent SO_3^{2-}

1

[12]

Q24.

(a) Hydrogen/ H_2 gas/bubbles

1

1.0 mol dm⁻³ HCl/ H^+

1

At 298K and 100kPa

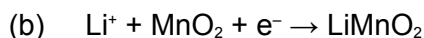
Allow 1 bar instead of 100 kPa

Do not allow 1 atm

1

Pt (electrode)

1

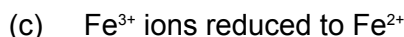


Ignore state symbols

1

-0.13(V)

1



Can score from equation/scheme

1

Because $E(\text{Fe}^{3+}/\text{Fe}^{2+}) > E(\text{H}^+/\text{H}_2)/E(\text{hydrogen})$

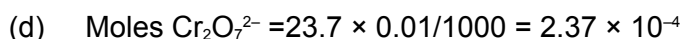
Allow $\text{emf}/E_{\text{cell}}$ +ve/0.77V

Allow Fe^{3+} better oxidising agent than H^+

Allow H_2 better reducing agent than Fe^{2+}

Only award this explanation mark if previous mark given

1



1

1 mol $\text{Cr}_2\text{O}_7^{2-}$ reacts with 6 mol Fe^{2+} so moles
 Fe^{2+} in $25 \text{ cm}^3 = 6 \times 2.37 \times 10^{-4} = 1.422 \times 10^{-3}$

1

M1 $\times 6$

Moles Fe^{2+} in $250 \text{ cm}^3 = 1.422 \times 10^{-2}$

M2 $\times 10$ or M4/10

1

Original moles $\text{Fe}^{2+} = \underline{10.00/277.9} = 0.0360$

Independent mark

1

Moles Fe^{2+} oxidised = $0.0360 - 0.0142 = 0.0218$

M4 – M3

1

% oxidised = $(0.0218 \times 100)/0.0360 = 60.5\%$

(M5 $\times 100$)/M4

Allow 60 to 61

Note Max 3 if mol ratio for M2 wrong

eg 1:5 gives 67.1%

1:1 gives 93.4%

Note also, 39.5% (39-40) scores M1, M2, M3 and M4 (4 marks)

Q25.

(a) 1.4 V

Allow + or –

1

(b) $2\text{NiO}(\text{OH}) + 2\text{H}_2\text{O} + \text{Cd} \rightarrow 2\text{Ni}(\text{OH})_2 + \text{Cd}(\text{OH})_2$ *Mark for species, Deduct a mark for additional species (eg OH-) but allow balance mark*

1

Balanced

If equation is reversed CE=0

1

(c) NiO(OH) or Ni(III) or nickel

1

+3

Allow conseq on wrong species

1

[5]

Q26.

(a) (i) HgO

1

(ii) $\text{Hg}^{2+} + 2\text{e}^- \rightarrow \text{Hg}$

1

(iii) $2\text{H}_2\text{O} + \text{SO}_2 \rightarrow \text{H}_2\text{SO}_4 + 2\text{e}^-$ etc

1

(iv) $\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$

1

(b) (i) Vanadium species: VO_2^+

1

Oxidation state: 5

1

Half-equation: $\text{V}^{2+} + 2\text{H}_2\text{O} \rightarrow \text{VO}_2^+ + 4\text{H}^+ + 3\text{e}^-$

1

(ii) Cell e.m.f 0.06 V

1

Change in e.m.f , Increases

1

More Fe^{3+} ions to accept electrons

		1
	$\text{Fe}^{3+}/\text{Fe}^{2+}$ electrode becomes more positive	1
(c)	(i) $2\text{H}_2 \rightarrow 4\text{H}^+ + 4\text{e}^-$	1
	$4\text{e}^- + \text{O}_2 + 2\text{H}_2\text{O} \rightarrow 4\text{OH}^-$	1
	Overall equation $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$	
	(ii) Unchanged	1
(d)	Economic disadvantage; Use of CH_4 or cost of producing or high temp	1
	Environmental disadvantage; Makes CO_2	1
(e)	Cost of manufacture of solar cells	1
		[17]

Q27.

(a)	(i) Fe^{2+}	1
	(ii) F_2O	1
	(iii) Fe^{2+}	1
	Cl^-	1
<i>Use list principle if more than two answers</i>		
(b)	(i) e.m.f. = $E(\text{rhs}) - E(\text{lhs})$	1
	$= 1.52 - 0.77 = 0.75$ (0.75 scores first mark also)	1
	(ii) $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$	1
	(iii) Decrease (Increase is CE, no further marks)	1

Equilibrium (or reaction) shifts to R

(or L if refers to half equation in table)

(or in favour of more Fe^{3+})

(or more Fe^{3+} formed)

(or more electrons formed)

1

Electrode potential (for $\text{Fe}^{3+}/\text{Fe}^{2+}$) less positive (or decreases)

1

[10]

Q28.

(a) most powerful reducing agent: Zn;

1

(b) (i) reducing species: Fe^{2+}

1

(ii) oxidising species: Cl_2 ;

1

(c) (i) standard electrode potential 1.25 V;

1

(ii) equation: $\text{Ti}^{3+} + 2 \text{Fe}^{2+} \rightarrow 2\text{Fe}^{3+} + \text{Ti}$ + balanced;

1

correct direction;

1

(d) (i) moles $\text{KMnO}_4 = 16.2 \times 0.0200 \times 10^{-3} = 3.24 \times 10^{-4}$;

1

moles $\text{H}_2\text{O}_2 = \text{Moles KMnO}_4 \times 5 / 2 = 8.10 \times 10^{-4}$;

1

8.10×10^{-4} moles H_2O_2 in 25 cm^3

$8.10 \times 10^{-4} \times 1000 / 25$ in $1000 \text{ cm}^3 = 0.0324 \text{ mol dm}^{-3}$;

1

hence $\text{g dm}^{-3} = \text{mol dm}^{-3} \times M_r = 0.0324 \times 34 = 1.10$;

(penalise use of an incorrect H_2O_2 to KMnO_4 ratio by two marks)

1

(ii) $PV = nRT$;

1

hence $V = nRT / P$

$= 8.10 \times 10^{-4} \times 8.31 \times 298 / 98000$;

1

$= 2.05 \times 10^{-5}$;

1

units m^3 ;

(mark consequentially to answers in (c)(i))

(allow correct answers with other units)
 (answers to (c)(i) and (ii) must be to 3 significant figures;
 penalise once only)

1

[14]

Q29.

(a) (i) 0.60 V

1

(ii) $\text{H}_2\text{O} + \text{H}_2\text{SO}_3 \rightarrow \text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^-$

1

(b) (i) $2\text{IO}_3^- + 2\text{H}^+ + 5\text{H}_2\text{O}_2 \rightarrow 5\text{O}_2 + \text{I}_2 + 6\text{H}_2\text{O}$ Species

1

Balanced

1

(ii) The concentration of the ions change or are no longer standard or the e.m.f is determined when no current flows

1

(iii) Unchanged

1

(iv) Increased

1

Equilibrium IO_3^-/I_2 displaced to the right

1

Electrons more readily accepted or more reduction occurs or electrode becomes more positive (Q o L)

1

(c) VO_2^+

1

5 or V

1

$\text{V}^{2+} + 2\text{H}_2\text{O} \rightarrow \text{VO}_2^+ + 4\text{H}^+ + 3\text{e}^-$

1

[12]

Q30.

D

[1]

Q31.

C

[1]

Q32.

B

[1]

Q33.

D

[1]

Q34.

- (a) *Hydrogen ion concentration:* 1.00 mol dm⁻³ (1)

Hydrogen gas pressure: 100 kPa (1)

2

- (b) *Explanation of change:* Equilibrium displaced to left (1)
to reduce constraint (1)

Change in electrode potential: Becomes negative or decreases (1)
allow more negative

3

- (c) (i) 0.43V (1)

- (ii) *Half-equation:* $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$ (1)

Overall equation: $2\text{BrO}_3^- + 10\text{Br}^- + 12\text{H}^+ \rightarrow 6\text{Br}_2 + 6\text{H}_2\text{O}$ (2)
or $\text{BrO}_3^- + 5\text{Br}^- + 6\text{H}^+ \rightarrow 3\text{Br}_2 + 3\text{H}_2\text{O}$

species (1)
balanced (1)

4

[9]

Q35.

C

[1]

Examiner reports

Q1.

- (a) Although 73.4% of students were able to answer correctly, some stated that the circuit was completed by the flow of electrons.
- (b) Nearly two-thirds of the students (65.6%) could calculate the EMF of the cell.
- (c) Students knew that the concentrations of copper sulfate were different but only just over 20% of them could state that the left-hand electrode's concentration needed to be 1 mol dm^{-3} .
- (d) Many students did not give state symbols and a few gave extra phase boundaries and platinum electrodes.
- (e) Only the best students (15.5%) scored both marks. Many students could state that the concentration of copper ions would increase but then thought that the reagents would run out.

Q2.

- (a) Only 11.3% of students scored both marks here; sometimes Fe^{2+} was chosen but students could not explain their choice. The most common incorrect answer was Zn^{2+} .
- (b) Just over three-quarters of students could give the oxidation state.
- (c) The structure was often given, incorrectly, as a mirror image and the type of isomerism as optical; only 37.2% of students scored both marks.
- (d) This was poorly answered, with only just over a third of students giving the correct equation.
- (e) Over half of the students (55.6%) gave the correct equations; incorrect equations included charged vanadium species. Surprisingly, 12% of students were unable to make any attempt at this question.

Q3.

Just over two-thirds of students arrived at the correct answer. D (followed by A) was the commonest incorrect response.

Q4.

- (a) The oxidation states were well known, with 92.2% of students gaining both marks; +7 was the commonest wrong answer for the nitrate ion.
- (b) Just under half of students could identify the weakest reducing agent.
- (c) "Nitrate ions" was a common insufficient answer, and sulfuric acid a common incorrect answer. Only the very best students scored mark point two; the common incorrect answer was that the $E^\ominus$ for NO_3^- was more positive than the $E^\ominus$ for copper. The equation was often unbalanced, but the cell emf was generally calculated correctly.

Q8.

Part (a) posed a problem for some students because of the language they used – ‘most negative’ is more accurate than ‘lowest’. Performance in (b) was good but in (c) credit was too often given for vague answers and for expanding on non-standard conditions for both marks. The sacrificial anode questions in (d) were done well by the abler students but resulted in some vagueness from others including a restatement of the information in the stem of the question.

Q9.

The correct answer was given by the majority of the students in part (a)(i). Part (a)(ii) was well done although a few students could not state the correct units of concentration. In part (b) the two half-cells were generally identified but the cell was often reversed as was the equation given. In part (c) the correct cobalt species was generally identified as the oxidising agent but the equation was often not balanced and the oxidation product given as Co^{2+} or H^+ . Students found part (d) challenging. Very few gave $E^\ominus$ with the correct terminology; very often this was given the wrong way round e.g. $E^\ominus \text{Au}/\text{Au}^+$. Most students thought, incorrectly, that the reaction was between gold and water. The $E^\ominus$ – values should have indicated to them that a redox reaction between gold and water is not feasible.

Q10.

Most students gave a correct answer to part (a) but some continued to confuse an oxidising agent with a Lewis acid and incorrectly stated that an oxidising agent is an electron pair acceptor. Inaccurate terminology continues to disadvantage students when they are asked to use electrode potential data to explain the action of oxidising and / or reducing agents. For example, in part (b), the iron (II) ion was the weakest oxidising agent in the Table because the $E^\ominus$ value for the $\text{Fe}^{2+} / \text{Fe}$ electrode is the least positive (or most negative). It is not precise enough to refer to the electrode potential of iron. In a similar way when gold (I) ions are added to water, the water is oxidised because the electrode potential of the Au^+ / Au electrode is more positive than that of the $\text{O}_2 / \text{H}_2\text{O}$ electrode. It is correct in this context to say that Au^+ ions are a better oxidising agent than O_2 but it is not correct to say that Au^+ ions are a better oxidising agent than water. Parts (c) and (e)(i) were answered well by many students. Many answers to part (e)(ii) stated incorrectly that chloride ions would react with iron (II) ions. In part (f) similar imprecise use of the correct $E^\ominus$ notation led to no marks being awarded for the use of data.

Q11.

Few correct answers were seen to part (a). Some students did not comply with the convention of placing the hydrogen electrode on the left when measuring the standard electrode potential. Some failed to include all the phase boundaries, putting a comma between H_2 and H^+ instead of a solidus, whilst other students included too many phase boundaries at the extremities or an incorrect, additional platinum connected to iron. Parts (b) and (c) were answered well. All the elements were chosen by students in part (d)(i) and where the correct answer of chlorine was selected, the oxidation states in part (d)(ii) and the equation in part (e) were usually correct. It was pleasing to find that answers to parts (f)(i) and (f)(ii) were usually correct. Most students identified the positive electrode correctly in part (f)(iii) but only the highest-scoring students used an appropriate electrode potential from the table to explain their answer.

Q12.

- (a) The diagrams given for the half-cell in this part were occasionally poorly labelled and, too often, leniently marked. Full names or formulae of suitable iron(II) and iron(III) compounds must appear when the term ‘reagent’ is requested. In a few

cases, students giving iron(III) sulfate as one reagent failed to realise that 1.0 mol dm^{-3} solutions of this would give double the ion concentration required. When only one reagent was given, this was penalised but the concentration mark could still be awarded and this was sometimes harshly treated by some teachers.

- (b) This part was well understood by many but the use of 'movement of electrons' appeared occasionally and was not always penalised.

Q13.

Lots of correct answers were seen for in part (a) and (b) but part (c) caused many more problems with sketchy answers sometimes being credited with full marks. It was not clearly understood that a method for determining the point at which the copper(II) concentration was 0.5 mol dm^{-3} was needed. Calorimetry was a common error. A novel approach was to wait until the cell potential had fallen to half its original value but this is incorrect chemistry and so should not have been credited. This is a good example of where reference to the Assessment Adviser would be helpful.

Q14.

This was a testing question and in common with redox questions in previous years, performance by students was very mixed. In part (a), over 30% of students answered incorrectly that the salt bridge allowed electrons to flow through it. In part (b), students were expected to use their knowledge of copper chemistry to suggest that the chloride ions from the potassium chloride would react with the aqueous copper ions but only a very small percentage of students answered this correctly. Part (c) also proved to be a very demanding question. Only the best students gave a correct answer involving the higher concentration of copper ions in the left hand electrode. It was surprising that many students appeared to consider that the 0.2 mol dm^{-3} solution of CuSO_4 was more concentrated than the 1.0 mol dm^{-3} solution. Part (d) was very discriminating. The great majority of students answered incorrectly by stating that one of the reagents would be used up or had reacted completely. Answers to part (e) were much better. When students are asked to write an equation for an overall reaction, they should take care to cancel out identical species that appear on both sides of the final equation. In part (e)(ii), many students lost a mark because they failed to cancel out a lithium ion that appeared on both sides of the equation. In part (e)(iii), it was surprising to note that a considerable number of students thought that carbon dioxide would be produced by a reaction with an imaginary carbon electrode in the lithium cell.

Q15.

- (a) This part was usually well answered. The usual mistakes were to make some reference to impurities without specifying aluminium oxide or contradicting an answer by giving an incorrect formula.
- (b) Answers to this part were variable. Students familiar with the standard hydrogen electrode gave good answers, although many omitted to name the acid used.
- (c) In this part, most students realised that a reaction would occur but gave impossible products and lost the mark. The consequences of reaction were not well understood. A number of students gave answers implying that a salt bridge would not be necessary or describing details of one specific use of a standard hydrogen electrode, which did not relate to the question being asked.

Q16.

This proved to be a difficult question. In part (a), answers were often very scrappy and disappointing. Diagrams showed a half cell that had no chance of working. The salt bridge, if present rarely dipped into a solution and the electrode material was often iron instead of the correct material, platinum. Only very good students showed a solution containing iron(II) and iron(III) ions and only the very highest-scoring students stated clearly that all aqueous species should have a concentration of 1 mol dm^{-3} . Answers to part (b) were more encouraging but only the better students realised that no platinum electrode are necessary with copper and iron electrodes. In parts (c) and (d), most students were unable to use data from the table correctly. For instance, it is correct to state that $E^\ominus \text{Au}^+/\text{Au} > E^\ominus \text{O}_2/\text{H}_2\text{O}$ but not correct to state that $E^\ominus \text{Au}^+/\text{Au} > E^\ominus \text{H}_2\text{O}/\text{O}_2$ or that gold ions are better oxidising agents than water. In this reaction, water is the reducing agent. In part (d), only 6% of students recognised that silver ions would oxidise iron to iron(II) ions and then go on to oxidise iron(II) ions to iron(III) ions.

Q17.

Part (a) was answered well. The most common error was to give the meaning of oxidation rather than an oxidising agent. Answers to part (b) were very often incorrect. All of the species from the table were seen in students' answers. The most common wrong answer was probably 'cadmium'. It still seems to be the case that many students do not understand that when redox half-equations are presented in their standard form with electrons on the left, the species on the left are oxidising agents and the species on the right are reducing agents. Parts (c)(i) and (ii) were answered correctly by most students. Part (c)(iii) proved to be more testing and only the best students recognised that the porous partition was acting in the same way as a salt bridge by permitting the transport of ions. Parts (c)(iv) and (v) were answered quite well although in part (c)(v) it was not always realised that the zinc container reacts to form zinc ions whilst the cell is in operation.

Part (d)(i) proved to be a discriminating question. An oxidation state of +2 for nickel in $\text{NiO}(\text{OH})$ was a common error and the equation was often unbalanced or not simplified by removing hydroxide ions. Part (d)(ii) was almost always answered correctly. In comparison, parts (e)(i) and (ii) were not answered well, usually because of a lack of attention to detail. Instead of using ethanol, a common mistake was to use CH_3OH . In part (e)(i), the equation for the combustion of ethanol was often unbalanced because students seemed to fail to take into account the oxygen in the ethanol. In part (e)(ii), despite being asked to write a half-equation, the electrons were often omitted or the numbers of hydrogen ions and electrons were wrong. Answers to part (e)(iii) were usually correct. In part (e)(iv), it seems that many students are still not clear about the meaning of the term carbon-neutral. The term is too often thought to refer just to carbon rather than to carbon dioxide in the atmosphere. Also, the majority of students did not make it clear that the carbon dioxide, released into the atmosphere by the processes that manufacture and use ethanol, can be balanced by the carbon dioxide taken out of the atmosphere by photosynthesis in plants.

Q18.

- (a) Students usually struggle with questions where data from a table of $E^\ominus$ values is used to explain a chemical reaction, and this part was no exception. Very few answers worth both marks were seen. Many students could not express themselves clearly while many others confused chlorine and the chloride ion. Statements such as 'manganate(VII) ions have a more positive electrode potential than chloride ions' and 'chlorine will oxidise manganese(II) ions' were very common and were often wrongly credited with a mark. In this question, students frequently ruined a promising answer by including some incorrect chemistry. Examples included 'manganate(VII) ions have a more positive electrode potential so will reduce chloride ions to chlorine' and 'hydrochloric acid was reduced by manganate(VII) ions'. It was

difficult to see how an answer such as 'HCl is not electronegative enough to acidify KMnO_4 ' could have been given any marks but it was.

- (b) The calculation in this part discriminated well. There were a lot of complete answers but many students, as expected, forgot that sulfuric acid is a dibasic acid.
- (c)
 - (i) The equation in this part must have been unfamiliar to most students but it proved straightforward.
 - (ii) In this part, most students appreciated that the presence of a brown precipitate would obscure the end-point, but the effect on the titre was not widely understood. Many students made a vague statement that the titre would be changed, but didn't explain the change itself.
- (d) Most students appreciated the effect of a large volume of water in this part, but many added that manganate(VII) was a weak oxidising agent in solution so lost the mark because of the chemical contradiction.

Q19.

In part (a), only about 20% of students scored all three marks. In order to gain full marks it was necessary to identify solution **A** as, for example, hydrochloric acid not just as H^+ ions. Part (b) was answered well. Answers to part (c) very rarely scored all three marks. The most common error was to suggest that the salt bridge allowed transfer of electrons rather than movement of ions. Answers to parts (d), (e)(i) and (e)(ii) were pleasing and about half of all students gained the mark in each case.

Q20.

Part (a) was answered well. In part (b), answers were disappointing. Only about 20% of students used data from the table correctly. An effective way of using the data would have been to state that the electrode potential for the chlorine/chloride half equation is more positive than for the oxygen/water half equation. However, in many cases the species involved in the half equation were not mentioned or were in the wrong order. Answers to part (c) were often correct. Answers to part (d) were often disappointing and many students suggested that the oxidation state of hydrogen rather than oxygen was changing. Part (e) required similar understanding and use of electrode potentials as in part (b) and responses were similarly disappointing.

Q21.

As usual, this question that tested candidates' understanding of redox and cells proved to be difficult but it was pleasing to note that even the weaker candidates were able to gain a significant number of marks. Parts (a) and (b) were answered well. Many candidates were able to score two out of three marks for part (c). The mark lost was usually due to incorrect use of data from the table. Many candidates stated that the electrode potential values showed that fluorine is a better oxidising agent than water. In this case, water is the reducing agent and candidates should have stated that the data shows that fluorine is a better oxidising agent than oxygen. In part (d)(i), a majority of candidates scored one mark out of the two available; only the best candidates showed metal species on the outside of the cell and all the phase boundaries, correctly indicated. Parts (d)(ii) and (e)(i) were answered well but only about a third of candidates gave a correct answer to part (d)(iii). In part (e)(ii) marks were lost because candidates did not simplify the overall equation after combining the half equations or the equation was for discharge rather than for recharging. As expected, part (f) proved to be challenging, particularly part (f)(iii). Fuel cells were included in the Specification as an example of 'How Science Works'. Their value is that they can produce efficiently a steady current to drive electric motors because the

concentrations of the reactants remain constant when the fuel and the oxidant are supplied continuously.

Q22.

This Question proved very demanding as expected. Many candidates provided some $E^\ominus$ calculations without any explanation of the significance of the results, and could not be given any marks. Confusion between chlorine and chloride ion was very common, often rendering the answer meaningless. A fair number of candidates thought that the oxidation of chloride ions took place instead of the oxidation of iron(II) ions. Full marks were quite rare.

Q23.

Redox questions involving electrochemical cells are usually found to be very difficult. It was therefore very pleasing to find that this question was answered well in all parts except for parts (a)(iv) and (b)(i). In part (a)(iv) candidates were expected to recall from Unit 1 that lithium reacts with water but only about 20% were able to do this and these were not necessarily the best candidates. As expected, part (b)(i) proved to be discriminating. Although about one third of candidates scored both marks, others failed to include platinum or put the various ions in the wrong order.

Q24.

This was the most difficult question in the paper. In part (a) most candidates knew that a standard hydrogen electrode uses hydrogen gas and a platinum electrode but only a minority went on to state the correct and necessary conditions of concentration, temperature and pressure for operation of the electrode. In part (b) most candidates gave a correct value for the electrode potential of the positive electrode but very few candidates were able to use the cell representation to deduce a half-equation for the electrode reaction. Answers to part (c) were also disappointing. Despite the given electrode potentials and half-equations, most candidates could not predict and explain the redox reaction between hydrogen and iron(III) ions. Part (d) also proved to be a difficult question. Candidates were usually able to score three marks out of six for a correct calculation of the number of moles of dichromate(VI) ions, a calculation of the original number of moles of iron(II) sulfate and for recognising the involvement of a factor of ten. Only the very best candidates realised that the number of moles of iron ions was six times the number of moles of dichromate(VI) ions and that a subtraction was required to calculate the percentage of iron(II) ions that had been oxidised by air before the titration.

Q25.

This question was answered well except for part (b) where only the best candidates were able to write a balanced equation with the oxidising and reducing agent in the correct position.

Q27.

This question discriminated very effectively between candidates. In part (a) weak candidates were confused about reducing and oxidising agents. These agents were confused with the species that could be reduced and oxidised. Thus for part (a)(i) Fe^{3+} was a very common wrong answer and for part (a)(ii) the F^- ion was often suggested. Similarly in part (a)(iii) species were often incorrectly chosen from the left-hand side of the redox half-equations. Parts (b)(i) and (b)(ii) were usually correct but candidates were less successful with part (b) (iii). About half of all candidates gave the incorrect answer, an increase in the concentration of Fe^{3+} ions. For those candidates who gave a correct

change in concentration only the best gained two marks for the explanation. Most omitted to mention that the electrode potential of the $\text{Fe}^{3+}/\text{Fe}^{2+}$ electrode would be reduced.

Q28.

Although parts (a), (b) and (c) were rather better answered than similar questions on earlier papers, a significant number of candidates still have problems using redox data. The calculation in (d)(i) was well done but many candidates lost one mark as, having calculated the concentration of the hydrogen peroxide solution in mol dm^{-3} , they failed to convert this to g dm^{-3} as required by the question. Part(d)(ii) was marked consequentially to answers determined in (d)(i). Almost all candidates stated the ideal gas equation and used it correctly but only the best candidates gave correct units for the volume of oxygen evolved.

Q29.

Almost every candidate calculated the cell e.m.f correctly and gave a correct half-equation in answer to part (a). Many also gave a correct equation in part (b) (i) but candidates who failed to cancel out the excess hydrogen ions lost one mark. By contrast, few stated in (b) (ii) either that when current flows the concentration of the species change or that the e.m.f. is measured when no current flows. Most answered (b) (iii) correctly. The effect on the cell e.m.f of changing the concentration of IO_3^- ions, required in (b) (iv), was deduced correctly by the better candidates many of whom also made the link to the position of the half-cell equilibrium process. Only the best then linked this change to the value of the electrode potential and stated that this becomes more positive. Sadly, in part (c), many candidates failed to read the question carefully and gave answers based on the formation of the intermediate species, V^{3+} , rather than VO_2^+ , the species present at the end of the reaction with an excess of acidified potassium manganate(VII).

Q34.

Almost all candidates gave the correct hydrogen ion concentration and the correct hydrogen gas pressure in part (a). Most also explained correctly why the electrode potential of a hydrogen electrode changes when the hydrogen ion concentration is reduced. Using this information to deduce correctly that the electrode potential would become negative proved more difficult. Most candidates calculated the e.m.f of the cell in part (c)(i) and many also gave a correct half-equation and a correct overall equation in part (c)(ii).

Q1.

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

	$E^{\ominus} / \text{V}$
$\text{MnO}_4^- (\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+ 1.52
$\text{Cr}_2\text{O}_7^{2-} (\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$	+ 1.33
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq})$	+ 0.77
$\text{Cr}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Cr}^{2+}(\text{aq})$	- 0.41
$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn}(\text{s})$	- 0.76

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

- A $\text{Fe}^{2+}(\text{aq})$ can reduce acidified $\text{MnO}_4^- (\text{aq})$ to $\text{Mn}^{2+}(\text{aq})$
- B $\text{CrO}_7^{2-} (\text{aq})$ can oxidise acidified $\text{Fe}^{2+}(\text{aq})$ to $\text{Fe}^{3+}(\text{aq})$
- C $\text{Zn}(\text{s})$ can reduce acidified $\text{Cr}_2\text{O}_7^{2-} (\text{aq})$ to $\text{Cr}^{2+}(\text{aq})$
- D $\text{Fe}^{2+}(\text{aq})$ can reduce acidified $\text{Cr}^{3+}(\text{aq})$ to $\text{Cr}^{2+}(\text{aq})$

(Total 1 mark)

Q2.

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

	$E^{\ominus} / \text{V}$
$\text{Ce}^{4+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ce}^{3+}(\text{aq})$	+1.70
$\text{MnO}_4^- (\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+1.51
$\text{Cl}_2(\text{g}) + 2\text{e}^- \rightleftharpoons 2\text{Cl}^-(\text{aq})$	+1.36
$\text{VO}_2^+(\text{aq}) + 2\text{H}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{VO}^{2+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$	+1.00
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	+0.77
$\text{SO}_4^{2-}(\text{aq}) + 4\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{H}_2\text{SO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l})$	+0.17

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

(1)

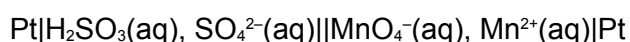
- (b) When the standard electrode potential for $\text{Fe}^{3+}(\text{aq}) / \text{Fe}^{2+}(\text{aq})$ is measured, a platinum electrode is required.

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

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

(3)

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



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 $\text{Fe}^{2+}(\text{aq})$ into $\text{Fe}^{3+}(\text{aq})$ but could not convert $\text{Mn}^{2+}(\text{aq})$ into $\text{MnO}_4^-(\text{aq})$?

(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).

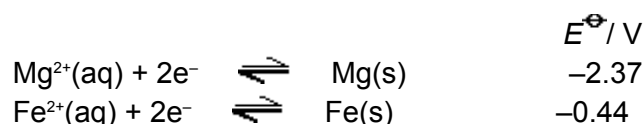
(2)

(Total 12 marks)

Q3.

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.





- (a) Calculate the e.m.f. of the cell represented by $\text{Mg}(\text{s})|\text{Mg}^{2+}(\text{aq})||\text{Fe}^{2+}(\text{aq})|\text{Fe}(\text{s})$ 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 $\text{Mg}(\text{s})|\text{Mg}^{2+}(\text{aq})||\text{Fe}^{2+}(\text{aq})|\text{Fe}(\text{s})$ 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 $\text{Pt}(\text{s})|\text{OH}^-(\text{aq})|\text{O}_2(\text{g})||\text{Fe}^{2+}(\text{aq})|\text{Fe}(\text{s})$ and use it to explain why iron corrodes when in contact with water which contains dissolved oxygen.

Cell e.m.f. _____

Explanation _____

_____ (2)

(Total 7 marks)

Mark schemes

Q1.

D

[1]

Q2.

(a) (Standard) hydrogen (electrode) (1)

1

(b) (i) To allow transfer of electrons / provide a reaction surface (1)

(ii) 298 K (1)

Both $\text{Fe}^{3+}(\text{aq})$ and $\text{Fe}^{2+}(\text{aq})$ have a concentration of

1 mol dm^{-3} (1) (QoL)

OR $[\text{H}^+] = 1 \text{ mol dm}^{-3}$

NOT zero current or 100 kPa

3

(c) +1.34 V (1)

$2 \text{MnO}_4^- + 5 \text{H}_2\text{SO}_3 \rightarrow 2 \text{Mn}^{2+} + 5 \text{SO}_4^{2-} + 3 \text{H}_2\text{O} + 4 \text{H}^+$

Correct species / order (1)

Balanced and cancelled (1)

Allow one for $2 \text{MnO}_4^- + 5 \text{H}_2\text{SO}_3 \rightarrow 2 \text{Mn}^{2+} + 5 \text{SO}_4^{2-}$

3

(d) (i) $\text{Ce}^{4+}(\text{aq})$ (1)

(ii) $\text{VO}_2^+(\text{aq})$ (1); Cl_2 (1)

Penalise additional answers to zero

3

(e) $\text{Pt} | \text{Fe}^{2+}(\text{aq}), \text{Fe}^{3+}(\text{aq}) || \text{Ce}^{4+}(\text{aq}), \text{Ce}^{3+}(\text{aq}) | \text{Pt}$

Correct species (1)

Correct order (1)

Deduct one mark for each error

2

[12]

Q3.

(a) Cell e.m.f.: 1.93 (v) CE if negative value given (1)

Half equation: $\text{Mg} \rightarrow \text{Mg}^{2+} + 2 \text{e}^-$ (1)

or $\rightleftharpoons$

Ignore state symbols

Mark on after an AE

2

(b) Change in e.m.f.: increases (1)

Mark on even if incorrect

Explanation: Equilibrium displaced to Mg^{2+} or to the left (1)

cell reaction or overall reaction goes to the right
Electrode is more negative or E decreases
or gives more electron
or forms more Mg^{2+} ions
Mark separately

3

(c) Cell e.m.f. : -0.84 (V) (1)

Explanation: Fe is giving electrons or forming Fe^{2+}
or reaction goes in the reverse direction (1)

Mark on after AE

N.B. In (a) and (c) mark on if no value given,
but CE in both (a) and (c) if e.m.f. = 0

2

[7]

Examiner reports

Q2.

There were many good answers to this question. Virtually all candidates were able to name the standard reference electrode in part (a). In part (b) only the best candidates stated the function of the platinum electrode correctly and many failed to state in section (ii) that, for standard conditions to apply, both $\text{Fe}^{3+}(\text{aq})$ and $\text{Fe}^{2+}(\text{aq})$ would need to be at a concentration of 1 mol dm^{-3} in the solution. Most deduced correctly the e.m.f. of the cell given in part (c).

Many scored one further mark in part (c) when they deduced the correct molar ratio of the reacting species in the equation but lost one mark when they failed to balance the equation correctly. In part (d) most candidates were able to deduce which of the species in the table of data was the strongest oxidising agent but many only gave one of the two correct answers in section (ii). A pleasing number of candidates gave a correct cell representation in part (e). Almost all were able to select the correct species for the cell but some then used the convention, as shown in part (c), incorrectly.

Q3.

This proved to be the most challenging question on the paper and candidates appeared unfamiliar with the conventional representation of cells. In part (a) many candidates were able to calculate the e.m.f. of the given cell but rather less gave a correct half-equation for the reaction occurring at the negative electrode. Some gave the overall equation for the reaction whilst others gave incorrect half-equations. Good candidates were able to deduce how a change in the concentration of Mg^{2+} ions would affect the cell e.m.f. and gave correct explanations for this change. Credit was given to candidates who correctly explained the effect of changing the concentration of Mg^{2+} ions even though they were unable to link this to the change in the e.m.f. of the cell. Part (c) was not well answered. Many candidates failed to use the conventional representation of cells to deduce that, in this case, the cell e.m.f. was negative and that the reaction would proceed spontaneously in the reverse direction with iron being oxidised to Fe^{2+} ions.