1 The table contains some standard electrode potential data.

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

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

....

....

(1)

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

Weakest oxidising agent ....

Explanation ....

....

....

(2)

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

State the conditions necessary when measuring this value.

Conventional representation ....

....

Conditions ....

....

....

....

(4)

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

State what you would observe.

Write an equation for the reaction that occurs.

Explanation ....

....

....

....

Observation ....

....

Equation ....

(4)

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

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

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

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

....

....

Answer ....

(1)

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

....

....

....

(1)

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

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....

....

(2)
(Total 15 marks)

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

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

- A sample of vase water was taken.
- An excess of potassium iodide solution was added to the sample.
- The chlorine in the sample oxidised the I^- ions to I_2
- The iodine was titrated with sodium thiosulfate ($Na_2S_2O_3$) solution.
- These steps were repeated using further samples taken from the vase water at hourly intervals.

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

Reason 1 ....

....

Reason 2 ....

....

(2)

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

....

....

....

(1)

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

....

....

(1)

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

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

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

....

....

....

(1)

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

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....

....

(1)
(Total 6 marks)

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

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

Explain the function of the salt bridge.

....

....

....

....

(2)

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

....

(1)

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

....

....

....

(1)

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

EMF = ... V

(1)

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

Suggest the major advantage of using the fuel cell.

....

....

....

(1)
(Total 6 marks)

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

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

Which statement is correct about the operation of the cell?

A Metallic copper is oxidised by Ag^+ ions.

B The silver electrode has a negative polarity.

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

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

(Total 1 mark)

5 The table below shows some standard electrode potential data.

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

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

....

(2)

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

....

....

(1)

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

Explain how you deduced your answer.

Species...

Explanation...

....

(2)

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

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

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

(1)

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

....

....

(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

....

....

....

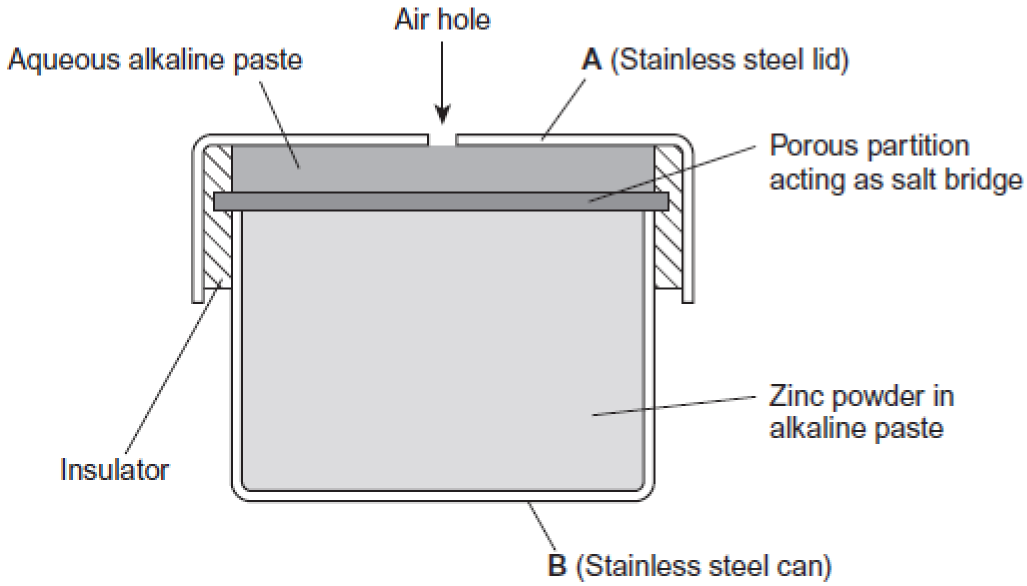
(2)

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

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

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

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



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

....

....

Answer = ...

(1)

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

....

....

....

(1)

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

Positive electrode ....

Reason ....

....

....

(2)

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

....

....

(1)
(Total 14 marks)

6 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)

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

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

(3)

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

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

Purpose of salt bridge ....

....

Requirement ....

....

(2)
(Total 5 marks)

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

(a) The conventional representation for the Daniell cell is

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

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

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

At Zn electrode ....

At Cu electrode ....

(2)

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

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

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(3)

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

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

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(3)
(Total 8 marks)

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

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

Explain why rechargeable cells are connected to these solar cells.

....

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(2)

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

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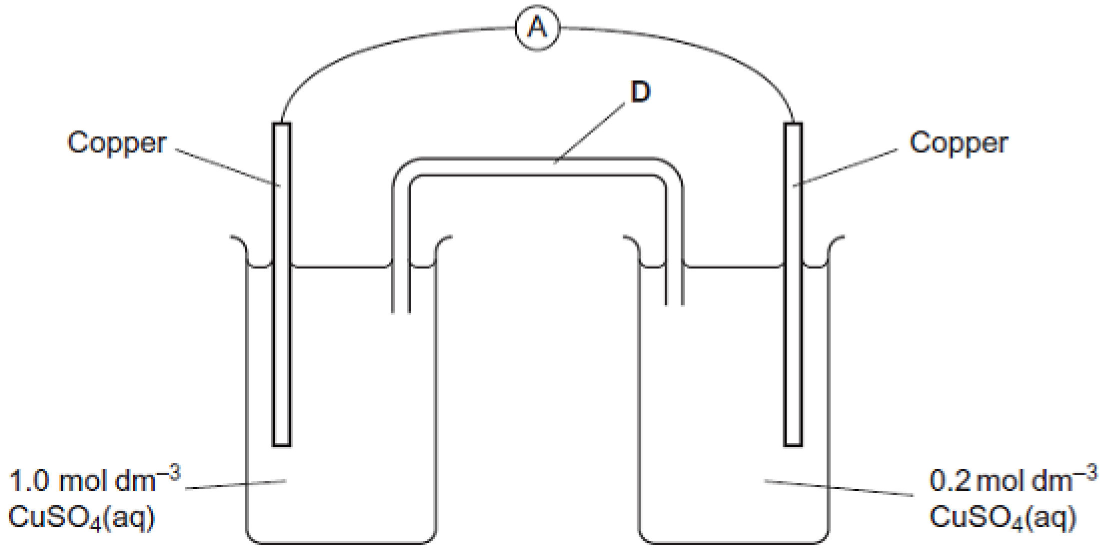
(1)
(Total 3 marks)

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



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

....

....

(1)

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

....

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(1)

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

Suggest why the electrons move in this direction.

....

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....

(2)

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

....

....

(1)

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

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

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

Electrode **B** is the negative electrode.

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

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

....

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(1)

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

....

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(2)

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

....

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(1)
(Total 9 marks)

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

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

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

Suggest the reason for this.

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(1)

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

(2)

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

....

....

....

....

(2)
(Total 5 marks)

12 This table shows some standard electrode potential data.

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

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

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

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

Conditions ....

....

....

(5)

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

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(4)

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

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

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(3)

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

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(3)
(Total 15 marks)

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

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

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

....

....

(1)

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

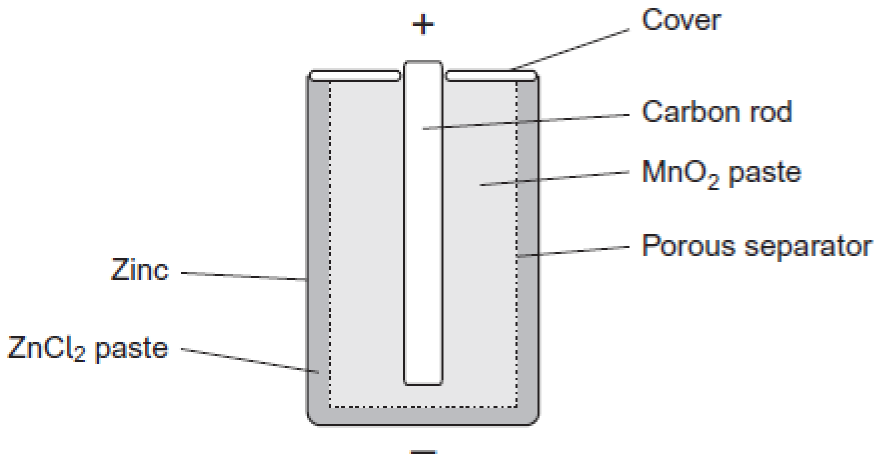
Weakest oxidising agent ....

Explanation ....

....

(2)

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



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

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....

....

(1)

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

....

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(1)

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

....

....

(1)

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

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(1)

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

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(1)

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

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

Oxidation state ....

Equation ....

....

....

(3)

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

....

....

(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.

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(1)

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

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(2)
(Total 17 marks)

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

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

....

....

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....

(2)

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

....

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....

(3)

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

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

....

....

(1)

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

....

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....

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(2)

(d) Potassium manganate(VII) is an oxidising agent.

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

....

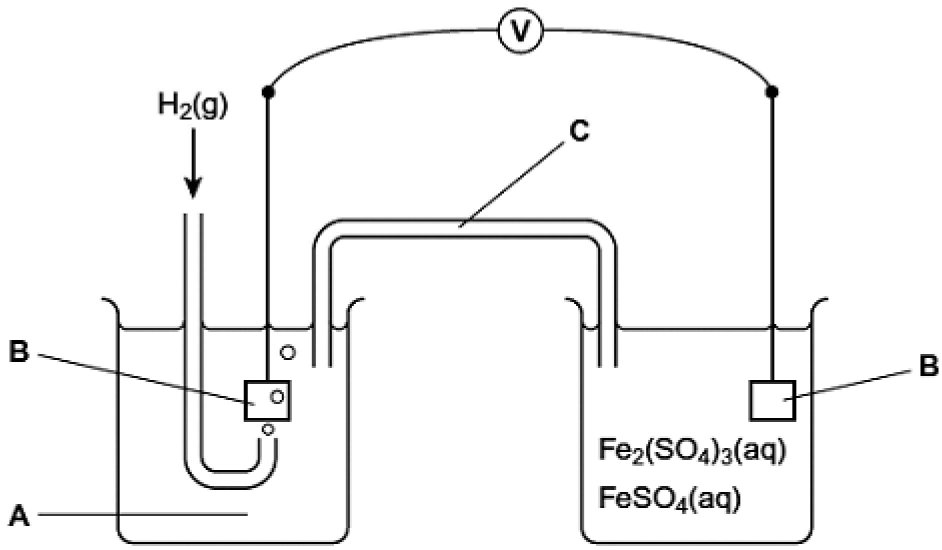
....

....

(1)

(Total 9 marks)

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



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

Solution **A** ....

Conditions ....

....

....

(3)

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

Material ....

Reason 1 ....

....

Reason 2 ....

....

(3)

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

Solution ....

Reason 1 ....

....

Reason 2 ....

....

(3)

(d) Write the conventional representation for this cell.

....

(1)

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

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

....

....

....

(1)

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

....

....

(1)
(Total 12 marks)

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

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

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

....

(1)

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

Explanation ....

....

....

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)

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

Half-equation	$E^{\ominus}$/ V
$Zn^{2+}(aq) + 2e^{-} \longrightarrow Zn(s)$	–0.76
$Ag_2O(s) + 2H^{+}(aq) + 2e– \longrightarrow 2Ag(s) + H_2O(l)$	+0.34
$O_2(g) + 4H^{+}(aq) + 4e^{-} \longrightarrow 2H_2O(l)$	+1.23
$F_2(g) + 2e^{-} \longrightarrow 2F^{-}(aq)$	+2.87

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

....

(1)

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

Explain how you deduced your answer.

Strongest reducing agent ....

Explanation ....

....

(2)

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

Write an equation for the reaction that occurs.

Explanation ....

....

....

Equation ....

....

(3)

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

(i) Give the conventional representation for this cell.

....

(2)

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

....

(1)

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

....

....

(1)

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

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

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

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

....

....

(1)

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

....

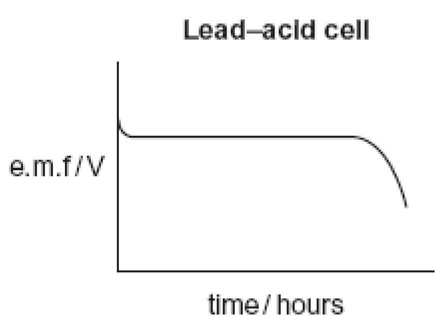
....

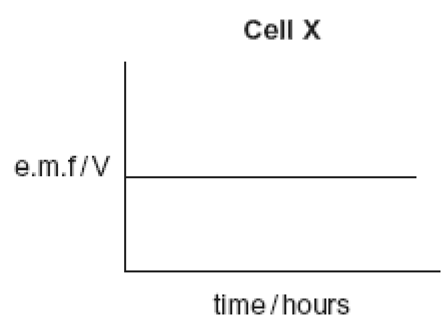
....

....

(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**.

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(Extra space) ....

....

(2)
(Total 17 marks)

18 The table below shows some standard electrode potentials.

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

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

....

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....

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(Total 2 marks)

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

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

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

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

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

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

Element ....

Oxidation state 1 ....

Oxidation state 2 ....

....

(3)

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

....

(1)

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

Property 1 ....

Property 2 ....

(2)

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

....

....

(1)

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

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

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

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

....

(2)

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

Overall equation ....

....

....

Oxidising agent ....

Reducing agent ....

(3)
(Total 12 marks)

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

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

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

(a) Describe a standard hydrogen electrode.

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(4)

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

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

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

Calculate the standard electrode potential of this positive electrode.

....

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....

....

(2)

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

....

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....

....

(2)

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

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

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(6)
(Total 14 marks)

21 Ethanedioic acid is an important industrial chemical with a number of uses.

Ethanedioate ions, $C_2O_4^{2-}$, act as bidentate ligands with transition metal ions.

(a) Write an equation for the ligand substitution reaction of an excess of ethanedioate ions with aqueous cobalt(II) ions.

....

(1)

(b) The table below shows some standard electrode potentials.

		$E^{\ominus}$ / V
$Fe^{3+}(aq) + e^-$	$\rightarrow Fe^{2+}(aq)$	+0.77
$2CO_2(g) + 2e^-$	$\rightarrow C_2O_4^{2-}(aq)$	–0.49

Use $E^{\ominus}$ values from the table to explain why an iron(III) complex is **not** formed when solutions containing ethanedioate ions and iron(III) ions are mixed.

....

....

....

....

(2)
(Total 3 marks)

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

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

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

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

....

(1)

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

....

....

....

(2)

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

Oxidising agent ....

Oxidation state ....

(2)
(Total 5 marks)

23 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*^o /V**
$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$	+1.23
$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$	+0.40
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$2H_2O(l) + 2e^- \rightarrow 2OH^-(aq) + H_2(g)$	– 0.83

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

....

(1)

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

....

(1)

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

....

(2)

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

....

....

....

....

(2)

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

....

....

(1)

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

....

(1)

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

....

....

(1)

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

....

(1)
(Total 10 marks)

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

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

Reaction **1** with HgO

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

Reaction **2** with chlorine

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

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

....

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

....

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

....

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

....

(4)

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

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

Each of the above can be reversed under suitable conditions

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

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

....

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

....

Half-equation ....

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

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

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

Cell e.m.f. ....

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

Explanation ....

....

(7)

(c) Consider the cell below

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

(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)

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

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

Each of the above can be reversed under suitable conditions.

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

....

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

....

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

....

(4)

(b) The cell represented below was set up.

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

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

....

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

Change in concentration ....

Explanation ....

....

....

(6)
(Total 10 marks)

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

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

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

....

(1)

(b) From the table above select

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

....

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

....

(2)

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

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

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

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

....

....

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

....

(3)

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

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

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

....

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....

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

....

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....

....

....

(8)
(Total 14 marks)

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

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

Each of the above can be reversed under suitable conditions.

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

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

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

....

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

....

(2)

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

Pt | H_2O_2(aq), O_2(g) || IO_3^- (aq), I_2(aq) |Pt

(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 V^{2+}(aq) ions. Use the data given in the table to determine the vanadium species present in the solution at the end of this reaction. State the oxidation state of vanadium in this species and write a half-equation for its formation from V^{2+}(aq).

Vanadium species present at end of reaction ....

Oxidation state of vanadium in final species ....

Half-equation ....

(3)
(Total 12 marks)

28 In this question consider the data below.

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

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

A 0.93 V

B 0.67 V

C −0.67 V

D −0.93 V

(Total 1 mark)

29 In this question consider the data below.

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

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

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

B increasing the surface area of the Pt electrode.

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

D decreasing the pressure of H_2(g).

(Total 1 mark)

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

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

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

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

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

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

B Only Cu(I) undergoes disproportionation.

C Only Hg(I) undergoes disproportionation.

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

(Total 1 mark)

31

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

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

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

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

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

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

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

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

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

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

(Total 1 mark)

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

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

Each of the above can be reversed under suitable conditions.

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

Hydrogen ion concentration ....

Hydrogen gas pressure ....

(2)

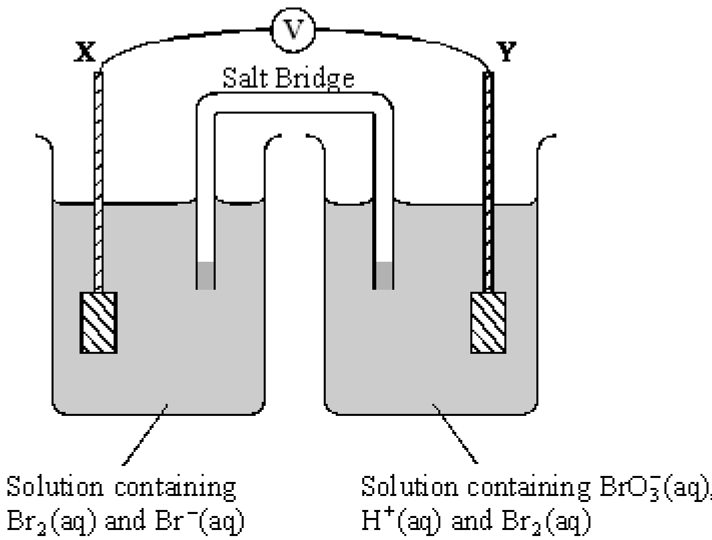
(b) The electrode potential of a hydrogen electrode changes when the hydrogen ion concentration is reduced. Explain, using Le Chatelier's principle, why this change occurs and state how the electrode potential of the hydrogen electrode changes.

Explanation of change ....

Change in electrode potential ....

(3)

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



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

....

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

Half-equation ....

Overall equation ....

....

(4)
(Total 9 marks)

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

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

The most powerful oxidising agent in the table is

A Mn^{2+}(aq)

B Zn(s)

C MnO_4^-(aq)

D Zn^{2+}(aq)

(Total 1 mark)

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

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

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

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

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

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

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

(Total 1 mark)

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

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

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

....

(1)

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

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

....

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

....

....

....

(3)

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

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

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

Cell e.m.f. …..

Equation ....

....

(3)

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

....

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

....

(3)

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

....

(2)
(Total 12 marks)

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

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

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

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

Cell e.m.f. ....

Half-equation ....

....

(2)

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

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

Explanation ....

....

(3)

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

Cell e.m.f. ....

Explanation ....

....

(2)
(Total 7 marks)

Mark schemes

1 (a) Electron acceptor / gains electrons

do not allow electron pair acceptor

1

(b) Fe^{2+} ions

1

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

Do not allow Fe / Fe^{2+}

Cannot score M2 if M1 incorrect

1

(c) $Pt|H_2|H^+||Ag^+|Ag$

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

1

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

1

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.

2

(d) E Au^+(/ Au) > E O_2 (/ H_2O) OR e.m.f. / *E*cell = 0.45 V

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

1

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

Allow water donates electrons to Au+

1

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

Penalise incorrect observations

1

$2Au^+ + H_2O \rightarrow 2Au + 2H^+ + \frac{1}{2}O_2$ OR $4Au^+ + 2H_2O \rightarrow 4Au + 4H^+ + O_2$

Allow multiples

1

(e) (i) 1.24 (V)

Do not allow −1.24

1

(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 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]

2 (a) (Biocide) reacts with bacteria / used up killing bacteria

Max two marks

Chlorine given off / evaporates

Do not allow "chlorine has reacted with water" alone.

Chlorine has reacted with water to form (HCl and) O_2

Do not allow products of HCl and HOCl alone

2

(b) the concentration of the remaining solution (after a sample has been removed) is unchanged.

1

(c) So that all chlorine was reacted / reduced

*Do **not** allow 'all of the iodide was oxidised'*

1

(d) The E^o value for the iodine half-equation is more positive than that for the thiosulfate

Allow = 0.45

Must refer to values

1

(e) $S_2O_3^{2-} + \frac{1}{2}I_2 \rightarrow I^- + \frac{1}{2}S_4O_6^{2-}$

Allow multiples

1

[6]

3 (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]

4 A

[1]

5 (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:

$Fe|Fe^{2+}||H^+|H_2|Pt$

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

(e) $4HOCl + 4H^+ + 4OH^- \rightarrow 2Cl_2 + O_2 + 6H_2O$

OR

$4HOCl \rightarrow 2Cl_2 + O_2 + 2H_2O$

Allow one mark for any incorrect equation that shows $HOCl \rightarrow Cl_2 + O_2$

Allow multiples

Ignore state symbols

Penalise one mark for uncancelled or uncombined species (eg $H_2O + H_2O$ instead of $2H_2O$)

2

(f) (i) e.m.f. = 0.40 – (–1.25) = <u>1.65</u> (V) / <u>+1.65</u> (V)

Allow –1.65 (V)

1

(ii) $2Zn + O_2 \rightarrow 2ZnO$

Allow multiples

Ignore state symbols

Do not allow uncancelled 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(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]

6 (a) (i) $H_2 + 2OH^- \rightarrow 2H_2O + 2e^-$ / $H_2 \rightarrow 2H^+ + 2e^-$

Any order

1

$O_2 + 4e^- + 2H_2O \rightarrow 4OH^-$ / $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$

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]

7

(a) Platinum electrode

1

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

Allow correct formulae for the iron compounds.

1

Concentrations of Fe(II) <u>and</u> Fe(III) ions are both 1 mol dm^{-3}

Ignore any references to temperature.

If eg $Fe_2(SO_4)_3$ used then concentration must be 0.5

1

(b) Purpose: Allow movement <u>of ions</u> 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]

8

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

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

1

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Ignore state symbols.

1

(b) Moles of copper(II) reacted = (100 / 1000) × 0.5 = 0.05

1

Moles of zinc reacted = 0.05

1

Mass of zinc lost = 0.05 × 65.4 = 3.27 g

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

1

(c) Allow cell to discharge until $[Cu^{2+}]$ 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]

9 (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]

10 (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) <u>Chloride</u> ions react with <u>copper ions</u> / <u>Cu^{2+}</u> ***OR*** $[CuCl_4]^{2-}$ formed

If incorrect chemistry, mark = 0

1

(c) The Cu^{2+} ions / $CuSO_4$ in the <u>left-hand</u> electrode more concentrated

Allow converse.

1

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

Allow <u>left-hand</u> electrode positive / <u>right-hand</u> electrode negative.

Also reduction at <u>left-hand</u> electrode / oxidation at <u>right-hand</u> electrode.

Also <u>left-hand</u> electrode has oxidising agent / <u>right-hand</u> 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) $LiMnO_2 \rightarrow Li + MnO_2$ 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

$LiMnO_2 \rightarrow Li + MnO_2$ scores 2

$Li^+ + LiMnO_2 \rightarrow Li^+ + Li + MnO_2$ scores 1

$Li + MnO_2 \rightarrow LiMnO_2$ 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]

11 (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 H_2 gas over a Pt electrode

Need H_2 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 Al^{3+}

Lose this mark if aluminium carbonate formed but mark on.

1

Reaction given (either equation or products specified)

OR H^+ / Al^{3+} concentrations change / cell e.m.f. altered

1

[5]

12 (a) Diagram of an Fe^{3+} / Fe^{2+} electrode that includes the following parts labelled:

Solution containing Fe^{2+} and Fe^{3+} 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

<u>all solutions</u> 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) $Cu^{2+} + Fe \rightarrow Cu + Fe^{2+}$

Ignore state symbols

1

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

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

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) $E^{\ominus}$ Au^{+}(/ Au) > $E^{\ominus}$ O_2 (/ H_2O)

Allow E cell / e.m.f. = 0.45 V

Allow 1.68 > 1.23

1

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

QoL

1

$2Au^{+} + H_2O \rightarrow 2Au + \frac{1}{2}O_2 + 2H^{+}$

Allow multiples

1

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

Allow E cell / e.m.f. = 1.24

Allow 0.80 > −0.44

1

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

Allow E cell / 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]

13 (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'

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) $2MnO_2 + 2H_2O + Zn \rightarrow 2MnO(OH) + 2OH^- + Zn^{2+}$

Ignore state symbols

e^- must be cancelled

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

1

(iii) Allows <u>ions</u> 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

$2Ni(OH)_2 + Cd(OH)_2 \rightarrow 2NiO(OH) + Cd + 2H_2O$

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) $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$

Allow C_2H_6O

1

(ii) $C_2H_5OH + 3H_2O \rightarrow 2CO_2 + 12H^+ + 12e^-$

Allow C_2H_6O

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]

14 (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 × 0.0200 × 8 / 1000 = 4.00 × 10^{-3}

1

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

Allow consequential marking on incorrect moles of H^+

1

Volume H_2SO_4 = 4.00 (cm^3) (2.00 × 10^{-3} × 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) $MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$

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]

15 (a) HCl 1.0 mol dm^{-3}

Allow H_2SO_4 0.5 mol dm^{-3}

Allow HNO_3 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, KNO_3, Na_2SO_4 etc NOT NH_4Cl

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) $Pt|H_2|H^+||Fe^{3+},Fe^{2+}|Pt$

Ignore state symbols

Order must be correct

| must be correct but allow | instead of , separating Fe^{3+} from Fe^{2+}

Allow , instead of | separating H_2 and H^+

1

(e) (i) $2Fe^{3+} + H_2 \longrightarrow 2Fe^{2+} + 2H^+$

Allow multiples

1

(ii) The <u>Fe</u>$^{3+}$ 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]

16 (a) H_2O_2

Ignore state symbols

1

(b) $E^\ominus$ Cl_2/Cl^- > $E^\ominus$ O_2/H_2O

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

Allow 1.36 > 1.23 / E cell = 0.13

1

$Cl_2 + H_2O \longrightarrow 2Cl^- + 1/2O_2 + 2H^+$

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

(d) <u>O</u> <u>(–1)</u> (in H_2O_2)

Must give oxidation state of O in H_2O_2= –1

1

Changes to <u>O(–2)</u> (in water)

Must give oxidation state of O in water = –2

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

1

(e) $E^\ominus$ H_2O_2/H_2O > $E^\ominus$ O_2/H_2O_2

Allow stated in words

Allow 1.77 > 0.68 / E cell = 1.09

1

$2H_2O_2 \longrightarrow O_2 + 2H_2O$

Allow multiples

H^+ and e^- must be cancelled

1

[8]

17 (a) loses electrons / donates electrons

penalise donates electron pair

1

(b) Zn

1

(most) negative E^o / lowest E^o / least positive

can only score M2 if M1 correct

do not allow e.m.f instead of E^o

1

(c) $\underline{E^o\ F_2}$ ($/F^-$) > $\underline{E^o\ O_2}$ ($/H_2O$)

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

$2F_2 + 2H_2O \rightarrow 4F^- + 4H^+ + O_2$

allow 4HF in equation

balanced equation scores M2 and M3

1

(d) (i) order correct Zn Zn^{2+} Ag_2O Ag or reverse of this order

ignore ss , H^+ and H_2O, no. of moles

1

all phase boundaries correct

allow $Zn|Zn^{2+}||Ag_2O,Ag$

or $Zn|Zn^{2+}||Ag_2O|H^+|Ag$ for M1 & M2

e.g. $Zn|Zn^{2+}||Ag_2O|Ag$ or $Ag|Ag_2O||Zn^{2+}|Zn$ 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) <u>Reaction(s</u>) 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]

18 MnO_4^- will oxidise the <u>chloride</u> ion / reaction of MnO_4^- and Cl^- feasible

Accept converse argument with $Cr_2O_7^{2-}$

Accept calculations of overall E° values.

1

Larger volume needed

1

[2]

19 (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) $Li \rightarrow Li^+ + e^-$

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) $Pt | SO_3^{2-}(aq), SO_4^{2-}(aq) || ClO_3^{-}(aq), Cl^{-}(aq) | Pt$

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

$Pt | Cl^-, ClO_3^- || SO_4^{2-}, SO_3^{2-} | Pt$

2

(ii) $ClO_3^- + 3SO_3^{2-} \rightarrow Cl^- + 3SO_4^{2-}$

1

Oxidising agent ClO_3^-

1

Reducing agent SO_3^{2-}

1

[12]

20 (a) Hydrogen/H_2 <u>gas/bubbles</u>

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

(b) $Li^+ + MnO_2 + e^- \rightarrow LiMnO_2$

Ignore state symbols

1

–0.13(V)

1

(c) Fe^{3+} ions reduced to Fe^{2+}

Can score from equation/scheme

1

Because $E(Fe^{3+}(/Fe^{2+})) > E(H^+/H_2)/E$(hydrogen)

Allow emf/E_{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

(d) Moles $Cr_2O_7^{2-}$ = <u>23.7 × 0.01/1000</u> = 2.37×10^{-4} **1**

1 mol $Cr_2O_7^{2-}$ reacts with 6 mol Fe^{2+} so moles Fe^{2+} in 25 cm^3 = $6 \times 2.37 \times 10^{-4} = 1.422 \times 10^{-3}$ **1**

M1 × 6

Moles Fe^{2+} in 250 cm^3 = 1.422×10^{-2}

M2 × 10 or M4/10 **1**

Original moles Fe^{2+} = <u>10.00/277.9</u> = 0.0360

Independent mark **1**

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

M4 – M3 **1**

% oxidised = (0.0218 × 100)/0.0360 = 60.5%

(M5 × 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]

21 (a) $3C_2O_4^{2-} + [Co(H_2O)_6]^{2+} \rightarrow [Co(C_2O_4)_3]^{4-} + 6H_2O$

Accept multiples.

Equation must have cobalt(II) hexaaqua ion. **1**

(b) Ethanedioate ion reduces iron(III) ion **or**
iron(III) ion oxidises ethanedioate ion

Allow answer using equations. **1**

$E^\ominus$ (CO_2 / $C_2O_4^{2-}$) more negative than $E^\ominus$ (Fe^{3+} / Fe^{2+}) **or**
$E^\ominus$ (Fe^{3+} / Fe^{2+}) > $E^\ominus$ (CO_2 / $C_2O_4^{2-}$)
or e.m.f. positive **or** cell voltage = +1.26 **1**

[3]

22 (a) 1.4 V

Allow + or – **1**

(b) $2NiO(OH) + 2H_2O + Cd \rightarrow 2Ni(OH)_2 + Cd(OH)_2$

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]

23 (a) By definition

allow 'set to this value'

1

(b) 1.23 V

Allow + or –

1

(c) $Pt|H_2(g)|OH^-(aq),H_2O(l)||O_2(g)|H_2O(l),OH^-(aq)|Pt$

H_2O not essential, allow reverse order

Correct but with Pt missing

1

Includes Pt with correct representation

1

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

And (2×) $2OH^- + H_2 \rightarrow 2H_2O + 2e^-$

1

$2H_2 + O_2 \rightarrow 2H_2O$

1

(e) Increases the surface area (so reaction faster)

1

(f) Overall reaction is the same ($2H_2 + O_2 \rightarrow 2H_2O$)

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]

24 (a) (i) HgO

1

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

1

(iii) $2H_2O + SO_2 \rightarrow H_2SO_4 + 2e^-$ etc

1

(iv) $Cl_2 + 2e^- \rightarrow 2Cl^-$

1

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

1

Oxidation state: 5

1

Half-equation: $V^{2+} + 2H_2O \rightarrow VO_2^+ + 4H^+ + 3e^-$

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

Fe^{3+}/Fe^{2+} electrode becomes more positive

1

(c) (i) $2H_2 \rightarrow 4H^+ + 4e^-$ **1**

$4e^- + O_2 + 2H_2O \rightarrow 4OH^-$ **1**

Overall equation $2H_2 + O_2 \rightarrow 2H_2O$

(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]

25 (a) (i) Fe^{2+} **1**

(ii) F_2O **1**

(iii) Fe^{2+} **1**

Cl^- **1**

Use list principle if more than two answers

(b) (i) e.m.f. = E(rhs) – E(lhs) **1**

= 1.52 – 0.77 = 0.75

(0.75 scores first mark also) **1**

(ii) $Fe^{2+} \rightarrow Fe^{3+} + e^-$ **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 Fe^{3+}/Fe^{2+}) less positive (or decreases)

1

[10]

26 (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: $Tl^{3+} + 2\ Fe^{2+} \rightarrow 2Fe^{3+} + Tl +$ balanced;

1

correct direction;

1

(d) (i) moles $KMnO_4 = 16.2 \times 0.0200 \times 10^{-3} = 3.24 \times 10^{-4}$;

1

moles H_2O_2 = Moles $KMnO_4 \times 5 / 2 = 8.10 \times {}^{-4}$;

1

8.10×10^{-4} moles H_2O_2 in 25 cm^3
$8.10 \times 10^{-4} \times 1000 / 25$ in 1000 cm^3 = 0.0324 mol dm^{-3};

1

hence g dm^{-3} = 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 × 10^{-4}× 8.31 × 298/98000; 1

= 2.05 × 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]

27 (a) (i) 0.60 V 1

(ii) $H_2O + H_2SO_3 \rightarrow SO_4^{2-} + 4H^+ + 2e^-$ 1

(b) (i) $2IO_3^- + 2H^+ 5H_2O_2 \rightarrow 5O_2 + I_2 + 6H_2O$ 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

$V^{2+} + 2H_2O \rightarrow VO_2^+ + 4H^+ \ 3e^-$

1

[12]

28 D

[1]

29 C

[1]

30 B

[1]

31 D

[1]

32 (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*: $2Br^- \rightarrow Br_2 + 2e^-$ **(1)**

Overall equation: $2BrO_3^- + 10Br^- + 12H^+ \rightarrow 6Br_2 + 6H_2O$ **(2)**
or $BrO_3^- + 5Br^- + 6H^+ \rightarrow 3Br_2 + 3H_2O$

*species **(1)***
*balanced **(1)***

4

[9]

33 C [1]

34 D [1]

35 (a) (Standard) hydrogen (electrode) **(1)** 1

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

(ii) 298 K **(1)**

Both F^{3+} (aq) and Fe^{2+} (aq) have a concentration of 1 mol dm^{-3} **(1)** (QoL)

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

NOT zero current or 100 kPa

3

(c) +1.34 V **(1)**

$2\ MnO_4^- + 5\ H_2SO_3 \rightarrow 2\ Mn^{2+} + 5\ SO_4^{2-} + 3\ H_2O + 4\ H^+$

Correct species / order **(1)**

Balanced and cancelled **(1)**

Allow one for $2\ MnO_4^- + 5\ H_2SO_3 \rightarrow 2\ Mn^{2+} + 5\ SO_4^{2-}$

3

(d) (i) Ce^{4+} (aq) **(1)**

(ii) VO_2^+ (aq) **(1)**; Cl_2 **(1)**

Penalise additional answers to zero

3

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

Correct species **(1)**

Correct order **(1)**

Deduct one mark for each error

2

[12]

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

Half equation: $Mg \rightarrow Mg^{2+} + 2\ e^-$ **(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]