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

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

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(2)

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

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(3)

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

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

....

....
(1)

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

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(2)

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

....

....

....
(1)
(Total 9 marks)

Q2. 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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(Total 2 marks)

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

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

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

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

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

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

Element ....

Oxidation state 1 ....

Oxidation state 2 ....

....

(3)

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

....

(1)

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

Property 1 ....

Property 2 ....

(2)

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

....

....

(1)

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

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

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

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

....

(2)

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

Overall equation ....

....

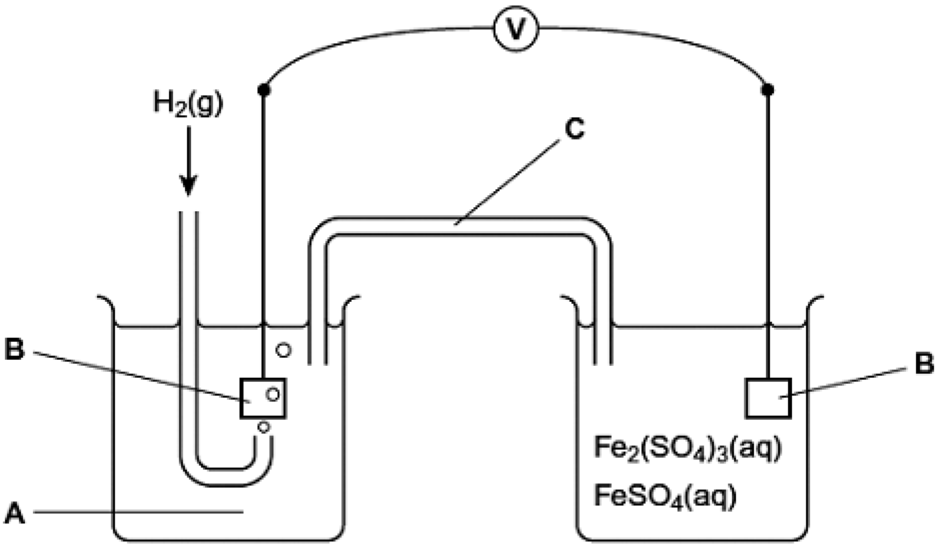
....

Oxidising agent ....

Reducing agent ....

(3)
(Total 12 marks)

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



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

Solution **A** ....

Conditions ....

....

....

(3)

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

Material ....

Reason 1 ....

....

Reason 2 ....

....

(3)

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

Solution ....

Reason 1 ....

....

Reason 2 ....

....

(3)

(d) Write the conventional representation for this cell.

....

(1)

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

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

....

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(1)

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

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(1)
(Total 12 marks)

Q5. Redox reactions occur in the discharge of all electrochemical cells. Some of these cells are of commercial value.
The table below shows some redox half-equations and standard electrode potentials.

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

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

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

....

(1)

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

Explain how you deduced your answer.

Strongest reducing agent ....

Explanation ....

....

(2)

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

Write an equation for the reaction that occurs.

Explanation ....

....

....

Equation ....

....

(3)

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

(i) Give the conventional representation for this cell.

....

(2)

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

....

(1)

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

....

....

(1)

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

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

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

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

....

....

(1)

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

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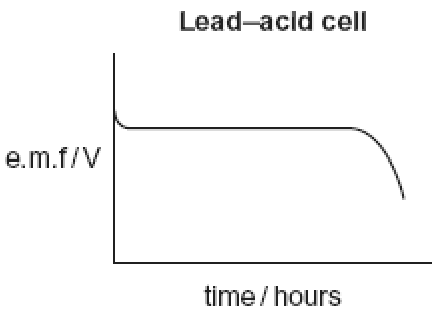
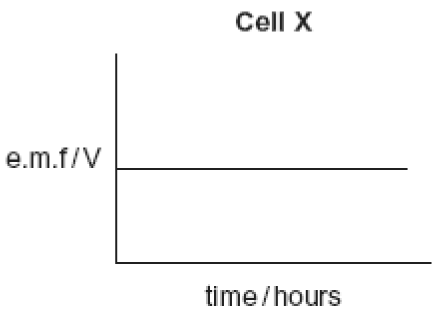
.... **(2)**

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





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

....

.... **(1)**

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

.... **(1)**

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

....

....

....

(Extra space) ....

.... **(2)**

(Total 17 marks)

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

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

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

....

(1)

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

Explanation ....

....

....

Equation ....

....

....

(2)

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

....

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(1)

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

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(2)

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

Explanation ....

....

Equation ....

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(2)
(Total 8 marks)