Q1. Use the standard electrode potential data given in the table below, where appropriate, to answer the questions which follow.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>$E^\circ/V$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$V^{3+}(aq) + e^- \rightarrow V^{2+}(aq)$</td>
<td>-0.26</td>
</tr>
<tr>
<td>$SO_3^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow H_2SO_3(aq) + H_2O$</td>
<td>+0.17</td>
</tr>
<tr>
<td>$VO^{2+}(aq) + 2H^+(aq) + e^- \rightarrow V^{3+}(aq) + H_2O(l)$</td>
<td>+0.34</td>
</tr>
<tr>
<td>$O_2(g) + 2H^+(aq) + 2e^- \rightarrow H_2O_2(aq)$</td>
<td>+0.68</td>
</tr>
<tr>
<td>$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$</td>
<td>+0.77</td>
</tr>
<tr>
<td>$VO_2^+(aq) + 2H^+(aq) + e^- \rightarrow VO^{2+}(aq) + H_2O(l)$</td>
<td>+1.00</td>
</tr>
<tr>
<td>$2IO_3^-(aq) + 12H^+(aq) + 10e^- \rightarrow I_2(aq) + 6H_2O(l)$</td>
<td>+1.19</td>
</tr>
<tr>
<td>$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$</td>
<td>+1.52</td>
</tr>
</tbody>
</table>

Each of the above can be reversed under suitable conditions.

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

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

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

\[
\text{Pt} | \text{H}_2\text{O}_2(\text{aq}), \text{O}_2(\text{g}) || \text{IO}^3(\text{aq}), \text{I}_2(\text{aq}) | \text{Pt}
\]

(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.

(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

(Total 12 marks)
Q2. Where appropriate, use the standard electrode potential data in the table below to answer the questions which follow.

<table>
<thead>
<tr>
<th>Species</th>
<th>Reduction reaction</th>
<th>$E^0/V$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{Zn}^{2+}(aq)$ + $2e^-$ → $\text{Zn}(s)$</td>
<td></td>
<td>-0.76</td>
</tr>
<tr>
<td>$\text{V}^{3+}(aq)$ + $e^-$ → $\text{V}^{2+}(aq)$</td>
<td></td>
<td>-0.26</td>
</tr>
<tr>
<td>$\text{SO}_4^{2-}(aq)$ + $2\text{H}^+(aq)$ + $2e^-$ → $\text{SO}_2^{3-}(aq)$ + $\text{H}_2\text{O}(l)$</td>
<td></td>
<td>+0.17</td>
</tr>
<tr>
<td>$\text{VO}^{2+}(aq)$ + $2\text{H}^+(aq)$ + $e^-$ → $\text{V}^{3+}(aq)$ + $\text{H}_2\text{O}(l)$</td>
<td></td>
<td>+0.34</td>
</tr>
<tr>
<td>$\text{Fe}^{3+}(aq)$ + $e^-$ → $\text{Fe}^{2+}(aq)$</td>
<td></td>
<td>+0.77</td>
</tr>
<tr>
<td>$\text{VO}_2^+(aq)$ + $2\text{H}^+(aq)$ + $e^-$ → $\text{VO}^{2+}(aq)$ + $\text{H}_2\text{O}(l)$</td>
<td></td>
<td>+1.00</td>
</tr>
<tr>
<td>$\text{Cl}_2(aq)$ + $2e^-$ → $2\text{Cl}^-(aq)$</td>
<td></td>
<td>+1.36</td>
</tr>
</tbody>
</table>

(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 $\text{VO}_2^+(aq)$ to $\text{VO}^{2+}(aq)$ but will not reduce $\text{VO}^{2+}(aq)$ to $\text{V}^{3+}(aq)$.

________________________________________________________________________

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

________________________________________________________________________

(2)

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

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

Cell e.m.f. = + 0.48 V

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

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

________________________________________________________________________

(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.

$$2\text{MnO}_4^- + 6\text{H}^+ + 5\text{H}_2\text{O}_2 \rightarrow 2\text{Mn}^{2+} + 8\text{H}_2\text{O} + 5\text{O}_2$$

(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.

(Total 14 marks)
Q3. (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

\[ \text{H}_2\text{O} + \text{SO}_2 + \text{HgO} \rightarrow \text{H}_2\text{SO}_4 + \text{Hg} \]

Reaction 2 with chlorine

\[ 2\text{H}_2\text{O} + \text{SO}_2 + \text{Cl}_2 \rightarrow \text{H}_2\text{SO}_4 + 2\text{HCl} \]

(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.

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

<table>
<thead>
<tr>
<th>Reaction</th>
<th>( \text{E} / \text{V} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{V}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{V}^{2+}(\text{aq}) )</td>
<td>-0.26</td>
</tr>
<tr>
<td>( \text{SO}_4^{2-}(\text{aq}) + 4\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2\text{SO}_3(\text{aq}) + \text{H}_2\text{O}(l) )</td>
<td>+0.17</td>
</tr>
<tr>
<td>( \text{VO}^{2+}(\text{aq}) + 2\text{H}^+(\text{aq}) + \text{e}^- \rightarrow \text{V}^{3+}(\text{aq}) + \text{H}_2\text{O}(l) )</td>
<td>+0.34</td>
</tr>
<tr>
<td>( \text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq}) )</td>
<td>+0.77</td>
</tr>
<tr>
<td>( \text{VO}^{2+}(\text{aq}) + 2\text{H}^+(\text{aq}) + \text{e}^- \rightarrow \text{VO}^{2+}(\text{aq}) + \text{H}_2\text{O}(l) )</td>
<td>+1.00</td>
</tr>
<tr>
<td>( \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}(l) )</td>
<td>+1.52</td>
</tr>
</tbody>
</table>

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.

\[ \text{Pt|H}_2\text{SO}_3(aq),\text{SO}_4^{2-}(aq),\text{H}^+(aq)||\text{Fe}^{3+}(aq),\text{Fe}^{2+}(aq)|\text{Pt} \]

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 ________________________________________________

(c) Consider the cell below

\[ \text{Pt|H}_2(\text{g})|\text{H}^+(aq)||\text{O}_2(\text{g})|\text{OH}^-(aq)|\text{Pt} \]

(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.

________________________________________________________

(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 ______________________________________

(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.

__________________________________________________________________

(Total 17 marks)
Q4. (a) Lithium ion cells are used to power cameras and mobile phones. A simplified representation of a cell is shown below.

Li | Li⁺ || Li⁺, CoO₂ | LiCoO₂ | 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.

\[ \text{Li}^+ + \text{CoO}_2 + e^- \rightarrow \text{Li}^+ [\text{CoO}_2]^- \]

(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 ____________________________________________

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

____________________________________________________________

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

Property 1 ____________________________________________
Property 2 ____________________________________________

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

____________________________________________________________

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

\[
\begin{align*}
\text{ClO}_3^- (aq) + 6H^+ (aq) + 6e^- & \rightarrow \text{Cl}^- (aq) + 3H_2O(l) \quad E^0 = +1.45 \text{ V} \\
\text{SO}_4^{2-} (aq) + 2H^+ (aq) + 2e^- & \rightarrow \text{SO}_3^{2-} (aq) + H_2O(l) \quad E^0 = +0.17 \text{ V}
\end{align*}
\]

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

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

Overall equation ____________________________________________________________________________

__________________________________________________________________________________________

__________________________________________________________________________________________

Oxidising agent ____________________________________________________________________________

Reducing agent ____________________________________________________________________________

(Total 12 marks)
Q5. 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.

<table>
<thead>
<tr>
<th>Half-equation</th>
<th>$E^\circ /V$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O} (l)$</td>
<td>+1.23</td>
</tr>
<tr>
<td>$\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(l) + 4\text{e}^- \rightarrow 4\text{OH}^- (\text{aq})$</td>
<td>+0.40</td>
</tr>
<tr>
<td>$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$</td>
<td>0.00</td>
</tr>
<tr>
<td>$2\text{H}_2\text{O}(l) + 2\text{e}^- \rightarrow 2\text{OH}^- (\text{aq}) + \text{H}_2(\text{g})$</td>
<td>−0.83</td>
</tr>
</tbody>
</table>

(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)
Q6. 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.

<table>
<thead>
<tr>
<th>Half-equation</th>
<th>$E^\circ/V$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{Cr}_2\text{O}_7^{2-}(aq) + 14\text{H}^+(aq) + 6e^- \rightarrow 2\text{Cr}^{3+}(aq) + 7\text{H}_2\text{O(l)}$</td>
<td>+1.33</td>
</tr>
<tr>
<td>$\text{Fe}^{3+}(aq) + e^- \rightarrow \text{Fe}^{2+}(aq)$</td>
<td>+0.77</td>
</tr>
<tr>
<td>$2\text{H}^+(aq) + 2e^- \rightarrow \text{H}_2(g)$</td>
<td>0.00</td>
</tr>
<tr>
<td>$\text{Fe}^{2+}(aq) + 2e^- \rightarrow \text{Fe(s)}$</td>
<td>−0.44</td>
</tr>
<tr>
<td>$\text{Li}^+(aq) + e^- \rightarrow \text{Li(s)}$</td>
<td>−3.04</td>
</tr>
</tbody>
</table>

(a) Describe a standard hydrogen electrode.

___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
___________________________________________________________________

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

$\text{Li(s)} | \text{Li}^+(aq) || \text{Li}^+(aq) | \text{MnO}_2(s) , \text{LiMnO}_2(s) | \text{Pt(s)}$

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.

___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
(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.

___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
___________________________________________________________________

(d) A solution of iron(II) sulfate was prepared by dissolving 10.00 g of FeSO$_4$.7H$_2$O ($M_r = 277.9$) 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.

___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
___________________________________________________________________

(Total 14 marks)
Q7. Use the data below, where appropriate, to answer the questions which follow.

### Standard electrode potentials

<table>
<thead>
<tr>
<th>Reaction</th>
<th>$E^\Theta/V$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$2\text{H}^+(\text{aq}) + 2e^- \rightarrow \text{H}_2(\text{g})$</td>
<td>0.00</td>
</tr>
<tr>
<td>$\text{Br}_2(\text{aq}) + 2e^- \rightarrow 2\text{Br}^-(\text{aq})$</td>
<td>+1.09</td>
</tr>
<tr>
<td>$2\text{BrO}_3^-(\text{aq}) + 12\text{H}^+(\text{aq}) + 10e^- \rightarrow \text{Br}_2(\text{aq}) + 6\text{H}_2\text{O}(\text{l})$</td>
<td>+1.52</td>
</tr>
</tbody>
</table>

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.

\[
\text{Hydrogen ion concentration} \quad \ldots \\
\text{Hydrogen gas pressure} \quad \ldots
\]

(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.

\[
\text{Explanation of change} \quad \ldots \\
\text{Change in electrode potential} \quad \ldots
\]
(c) A diagram of a cell using platinum electrodes X and Y is shown below.

Solution containing \( \text{Br}_2(\text{aq}) \) and \( \text{Br}^- (\text{aq}) \)  
Solution containing \( \text{BrO}_3^- (\text{aq}) \), \( \text{H}^+ (\text{aq}) \) and \( \text{Br}_2(\text{aq}) \)

(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

(Total 9 marks)
Q8. Use the data in the table below, where appropriate, to answer the questions which follow.

<table>
<thead>
<tr>
<th>Standard electrode potentials</th>
<th>( E^\circ / V )</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{Fe}^{3+}(aq) + e^- \rightarrow \text{Fe}^{2+}(aq) )</td>
<td>+0.77</td>
</tr>
<tr>
<td>( \text{Cl}_2(g) + 2e^- \rightarrow 2\text{Cl}^-(aq) )</td>
<td>+1.36</td>
</tr>
<tr>
<td>( 2\text{BrO}_3^- (aq) + 12\text{H}^+ (aq) + 10e^- \rightarrow \text{Br}_2(aq) + 6\text{H}_2\text{O}(l) )</td>
<td>+1.52</td>
</tr>
<tr>
<td>( \text{O}_3(g) + 2\text{H}^+(aq) + 2e^- \rightarrow \text{O}_2(g) + \text{H}_2\text{O}(l) )</td>
<td>+2.08</td>
</tr>
<tr>
<td>( \text{F}_2\text{O}(g) + 2\text{H}^+(aq) + 4e^- \rightarrow 2\text{F}^- (aq) + \text{H}_2\text{O}(l) )</td>
<td>+2.15</td>
</tr>
</tbody>
</table>

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 \( \text{BrO}_3^- (aq) \).

(b) The cell represented below was set up.

\[
\text{Pt} | \text{Fe}^{3+} (aq), \text{Fe}^{2+} (aq) || \text{BrO}_3^- (aq), \text{Br}_2(aq) | \text{Pt}
\]

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

\[
\text{Change in concentration}
\]

\[
\text{Explanation}
\]
Q9. 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.

\[ E^\circ / \text{V} \]

<table>
<thead>
<tr>
<th>Reaction</th>
<th>E^\circ</th>
<th>( \text{Mg}^{2+}(aq) + 2e^- \rightleftharpoons \text{Mg}(s) )</th>
<th>-2.37</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{Fe}^{2+}(aq) + 2e^- \rightleftharpoons \text{Fe}(s) )</td>
<td></td>
<td></td>
<td>-0.44</td>
</tr>
<tr>
<td>( \text{O}_2(g) + 2\text{H}_2\text{O}(l) + 4e^- \rightleftharpoons 4\text{OH}^-(aq) )</td>
<td></td>
<td></td>
<td>+0.40</td>
</tr>
</tbody>
</table>

(a) Calculate the e.m.f. of the cell represented by \( \text{Mg}(s)|\text{Mg}^{2+}(aq)||\text{Fe}^{2+}(aq)|\text{Fe}(s) \) 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 ______________________________________________________

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

Change in e.m.f. _____________________________________________________

Explanation ________________________________________________________

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

Cell e.m.f. _________________________________________________________

Explanation ______________________________________________________

(Total 7 marks)
Q10. Use the standard electrode potential data in the table below to answer the questions which follow.

\[
\begin{array}{ccc}
E/V & \\
\text{Ce}^{4+}(aq) + e^- & \rightarrow & \text{Ce}^{3+}(aq) & +1.70 \\
\text{MnO}_4^{-}(aq) + 8\text{H}^+(aq) + 5e^- & \rightarrow & \text{Mn}^{2+}(aq) + 4\text{H}_2\text{O}(l) & +1.51 \\
\text{Cl}_2(g) + 2e^- & \rightarrow & 2\text{Cl}^{-}(aq) & +1.36 \\
\text{VO}_2^+(aq) + 2\text{H}^+(aq) + e^- & \rightarrow & \text{VO}^{2+}(aq) + \text{H}_2\text{O}(l) & +1.00 \\
\text{Fe}^{3+}(aq) + e^- & \rightarrow & \text{Fe}^{2+}(aq) & +0.77 \\
\text{SO}_4^{2-}(aq) + 4\text{H}^+(aq) + 2e^- & \rightarrow & \text{H}_2\text{SO}_3(aq) + \text{H}_2\text{O}(l) & +0.17 \\
\end{array}
\]

(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.

\[
\text{Pt|H}_2\text{SO}_3(aq), \text{SO}_4^{2-}(aq)||\text{MnO}_4^{-}(aq), \text{Mn}^{2+}(aq)|\text{Pt}
\]

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

\[
\text{Cell e.m.f.}\]

\[
\text{Equation}\]

(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)?

(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).

(Total 12 marks)
Mark schemes

Q1.
(a)  (i)  0.60 V

(ii)  \[ H_2O + H_2SCO_3 \rightarrow SCO_4^2- + 4H^+ + 2e^- \]

(b)  (i)  \[ 2O_3^- + 2H^+ 5H_2O_2 \rightarrow 5O_2 + I_2 + 6H_2O \] Species

Balanced

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

(iii)  Unchanged

(iv)  Increased

Equilibrium \( IO_3^- / I_2 \) displaced to the right

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

(c)  \( VO_2^+ \)

5 or V

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

12

Q2.
(a)  most powerful reducing agent:  Zn;

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

(ii)  oxidising species:  \( Cl_2 \);

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

Electrochemical cell 2 SCT
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(ii) \[ \text{equation: } \text{Tl}^{3+} + 2 \text{Fe}^{2+} \rightarrow 2\text{Fe}^{3+} + \text{Tl} \quad \text{balanced;} \]

correct direction;

(d) (i) moles KMnO₄ = 16.2 \times 0.0200 \times 10^{-3} = 3.24 \times 10^{-4};

moles H₂O₂ = Moles KMnO₄ \times 5 / 2 = 8.10 \times 10^{-4};

8.10 \times 10^{-4} \text{ moles H}_2\text{O}_2 \text{ in } 25 \text{ cm}^3
8.10 \times 10^{-4} \times 1000 / 25 \text{ in } 1000 \text{ cm}^3 = 0.0324 \text{ mol dm}^{-3};

hence \( g \text{ dm}^{-3} = \text{mol dm}^{-3} \times M = 0.0324 \times 34 = 1.10; \)

(\text{penalise use of an incorrect H}_2\text{O}_2 \text{ to KMnO}_4 \text{ ratio by two marks})

(ii) \[ \text{PV} = nRT; \]

hence \( V = nRT / P \)
\[ = 8.10 \times 10^{-4} \times 8.31 \times 298 / 98000; \]
\[ = 2.05 \times 10^{-6}; \]

units m³;

(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)

Q3.

(a) (i) HgO

(ii) Hg²⁺ + 2e⁻ → Hg

(iii) 2H₂O + SO₂ → H₂SO₄ + 2e⁻ etc

(iv) Cl₂ + 2e⁻ → 2Cl⁻

(b) (i) Vanadium species: VO₂⁺

Oxidation state: 5

Half-equation: V^{2+} + 2H₂O → VO₂⁺ + 4H⁺ + 3e⁻
(ii) Cell e.m.f 0.06 V

Change in e.m.f, Increases

More Fe$^{3+}$ ions to accept electrons

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

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

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

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

(ii) Unchanged

(d) Economic disadvantage; Use of CH$_4$ or cost of producing or high temp

Environmental disadvantage; Makes CO$_2$

(e) Cost of manufacture of solar cells

Q4.

(a) (i) Co/Cobalt

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

(+) 4

(+) 3
 Allow 4 and 3 in either order

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

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

(iii) Platinum is a conductor

(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

(iv) \( \text{Li} \) reacts with water/forms lithium hydroxide
\( \text{Allow water breaks down (or is electrolysed) on re-charge} \)

(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
\( \text{Allow | in place of ‘,’ NOT ‘,‘ in place of |} \)
\( \text{Ignore H}^{+} \text{ and H}_2\text{O} \)
Deduct one mark for each mistake (e.g. Pt missed twice counts as two mistakes)
\( \text{Allow reverse order for whole cell} \)
\( \text{Pt} | \text{Cl}^{-}, \text{ClO}_3^{-} | \text{SO}_4^{2-}, \text{SO}_3^{2-} | \text{Pt} \)

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

Q5.

(a) By definition
\( \text{allow ‘set to this value’} \)

(b) 1.23 V
\( \text{Allow + or –} \)

(c) \( \text{Pt} | \text{H}_2(\text{g}) | \text{OH}^{-} \text{(aq)}, \text{H}_2\text{O(l)} | \text{O}_2(\text{g}) | \text{H}_2\text{O(l)}, \text{OH}^{-} \text{(aq)} | \text{Pt} \)
\( \text{H}_2\text{O not essential, allow reverse order} \)
Correct but with Pt missing
\( \text{Includes Pt with correct representation} \)

(d) Uses \( \text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^{-} \rightarrow 4\text{OH}^{-} \)
\( \text{And (2×) 2OH}^{-} + \text{H}_2 \rightarrow 2\text{H}_2\text{O} + 2\text{e}^{-} \)
\( 2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O} \)
(e) Increases the surface area (so reaction faster)  

(f) Overall reaction is the same (\(2H_2 + O_2 \rightarrow 2H_2O\))  
    \(Or \text{ shows e.m.f. is the same}\)

(g) Hydrogen and oxygen supplied continuously  
    
    OR

Can be operated without stopping to recharge  
    \(Or \text{ can be refuelled quickly}\)
    \(Allow \text{ any one mark}\)

(h) Hydrogen may need to be made using an energy source that is not ‘carbon neutral’

Q6.

(a) Hydrogen/H\(_2\) gas/bubbles  

1.0 mol dm\(^{-3}\) HCl/H\(^+\)  

At 298K and 100kPa  
    \(Allow 1 \text{ bar instead of 100 kPa}\)
    \(Do not allow 1 \text{ atm}\)

Pt (electrode)

(b) \(Li^+ + MnO_2 + e^- \rightarrow LiMnO_2\)  
    \(Ignore \text{ state symbols}\)

\(-0.13(V)\)

(c) \(Fe^{3+}\) ions reduced to \(Fe^{2+}\)  
    \(Can \text{ score from equation/scheme}\)

Because \(E(Fe^{3+}/Fe^{2+}) > E(H^+/H_2)/E(\text{hydrogen})\)
    \(Allow \text{ emf/E}_{cell} +ve/0.77V\)
    \(Allow \text{ Fe}^{3+} \text{ better oxidising agent than } H^+\)
    \(Allow \text{ H}_2 \text{ better reducing agent than } Fe^{2+}\)
    \(Only \text{ award this explanation mark if previous mark given}\)

(d) Moles \(Cr_2O_7^{2-} = 23.7 \times 0.01/1000 = 2.37 \times 10^{-4}\)
1 mol Cr$_2$O$_7^{2-}$ reacts with 6 mol Fe$^{2+}$ so moles
Fe$^{2+}$ in 25 cm$^3 = 6 \times 2.37 \times 10^{-4} = 1.422 \times 10^{-3}$

\[ M1 \times 6 \]

Moles Fe$^{2+}$ in 250 cm$^3 = 1.422 \times 10^{-2}$
\[ M2 \times 10 \text{ or } M4/10 \]

Original moles Fe$^{2+} = \frac{10.00}{277.9} = 0.0360$

Independent mark

Moles Fe$^{2+}$ oxidised = 0.0360 – 0.0142 = 0.0218
\[ M4 - M3 \]

% 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)

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**Q7.**

(a) *Hydrogen ion concentration:* 1.00 mol dm$^{-3}$ (1)

*Hydrogen gas pressure:* 100 kPa (1)

(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*

(c) (i) 0.43V (1)

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

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**Q8.**

Electrochemical cell 2 SCT
(a) (i) \( \text{Fe}^{2+} \)  
(ii) \( \text{Fe}_2\text{O} \)  
(iii) \( \text{Cl}^- \)  

*Use list principle if more than two answers*

(b) (i) e.m.f. = \( E(\text{rhs}) - E(\text{lhs}) \)  

\[
= 1.52 - 0.77 = 0.75
\]

(0.75 scores first mark also)  

(ii) \( \text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + e^- \)  

(iii) Decrease  

(*Increase is CE, no further marks*)  

Equilibrium (or reaction) shifts to R  

(or L if refers to half equation in table)  

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

(or more \( \text{Fe}^{3+} \) formed)  

(or more electrons formed)  

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

[10]

Q9.  

(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 \( \stackrel{\text{L}}{\longrightarrow} \)  

*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 \( \text{Mg}^{2+} \) 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 \( \text{Mg}^{2+} \) ions*  

*Mark separately*
Q10.

(a) (Standard) hydrogen (electrode) (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

(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 $SO_4^{2-}$

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

(ii) $VO_2^+$ (aq) (1); $Cl_2$ (1)
Penalise additional answers to zero

(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