MEGA LECTURE

A LEVEL CHEMISTRY

TOPIC 13 – ELECTROCHEMISTRY

ASSESSED HOMEWORK

Answer all questions

Max 80 marks

Name	……………………………………………………………..		
Mark	……../80	……….%	Grade ………

1. 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)

2. 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
$MnO_4^-(aq) + 8H^+(aq) + 5e^-$	⇌	$Mn^{2+}(aq) + 4H_2O(l)$	+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)

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

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

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....

....
....

....
....

(2)

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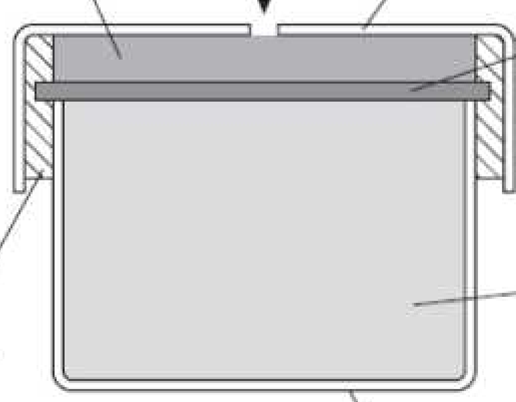
(f) The table is repeated to help you answer this question.

	$E^{\ominus}$ / 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)

4. 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.

....

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.... **(2)**

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

Explanation ....

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....

Equation ....
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....
....

(2)
(Total 8 marks)

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

Standard electrode potentials

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

6. (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.

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

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

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

Half-equation	***E*°/V**
$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$	+1.23

$O_2(g) + 2H_2O(l) + 4e^-$	→	$4OH^-(aq)$	+0.40
$2H^+(aq) + 2e^-$	→	$H_2(g)$	0.00
$2H_2O(l) + 2e^-$		$2OH^-(aq) + H_2(g)$	– 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.

....
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....

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....

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....

(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)

8.

$Cr_2O_7^{2-}$(aq) + 14H^+(aq) + 6e^-	→	2Cr^{3+}(aq) + 7H_2O(l)	$E^{\ominus}$ = +1.33 V
Br_2(aq) + 2e^-	→	2Br^-(aq)	$E^{\ominus}$ = +1.09 V
Fe^{3+}(aq) + e^-	→	Fe^{2+}(aq)	$E^{\ominus}$ = +0.77 V
VO^{2+}(aq) + 2H^+(aq) + e^-	→	V^{3+}(aq) + H_2O(l)	$E^{\ominus}$ = +0.34 V
SO_4^{2-}(aq) + 4H^+(aq) + 2e^-	→	H_2SO_3(aq) + H_2O(l)	$E^{\ominus}$ = +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)

9. 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)