HOLYDAY PACKAGE
CHEMISTRY S6 MCB1,2
GS ST PHILIPPE NERI

UNIT 12. ELECTROCHEMICAL CELLS & APPLICATIONS

PART 1.

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

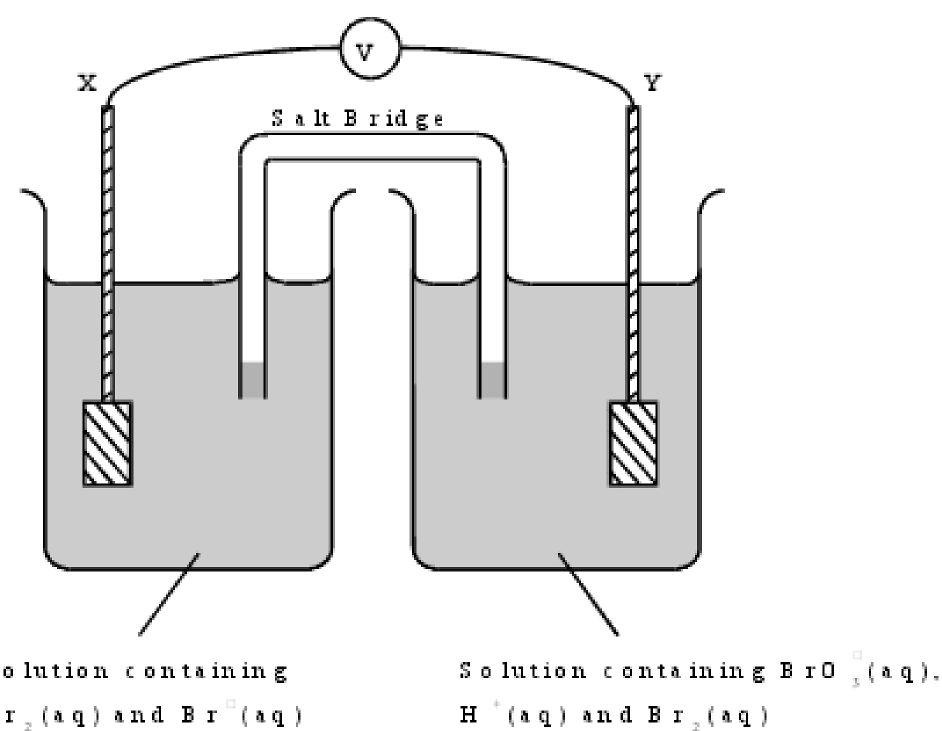
Standard electrode potentials	$E^{\ominus}$/V
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$Br_2(aq) + 2e^- \rightarrow 2Br^-(aq)$	+1.09
$2BrO_3^-(aq) + 12H^+(aq) + 10e^- \rightarrow Br_2(aq) + 6H_2O(l)$	+1.52

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.

(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.

(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.

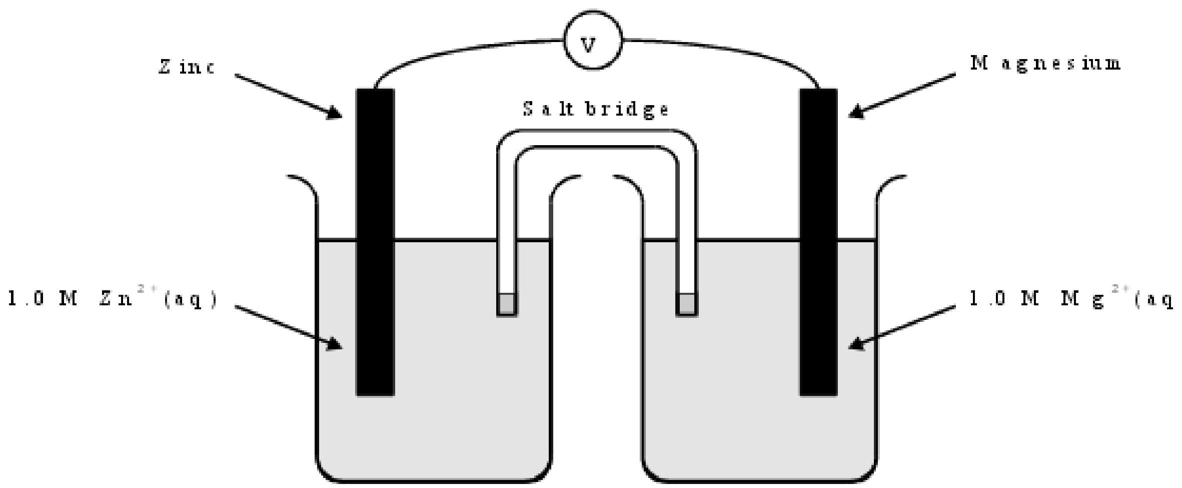
(ii) Write a half-equation for the reaction occurring at electrode **X** and an overall equation for the cell reaction which occurs when electrodes **X** and **Y** are connected.

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

Standard electrode potentials in acid solution	$E^{\ominus}$/V
$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$	–2.37
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	–0.76
$Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq)$	+0.15
$VO^{2+}(aq) + 2H^+(aq) + e^- \rightarrow V^{3+}(aq) + H_2O(l)$	+0.34
$VO_2^+(aq) + 2H^+(aq) + e^- \rightarrow VO^{2+}(aq) + H_2O(l)$	+1.02
$Ce^{4+}(aq) + e^- \rightarrow Ce^{3+}(aq)$	+1.70

(a) Give the components of the standard reference electrode used in determining the standard electrode potentials above. State the conditions under which this standard reference electrode has a potential of 0.00 V.

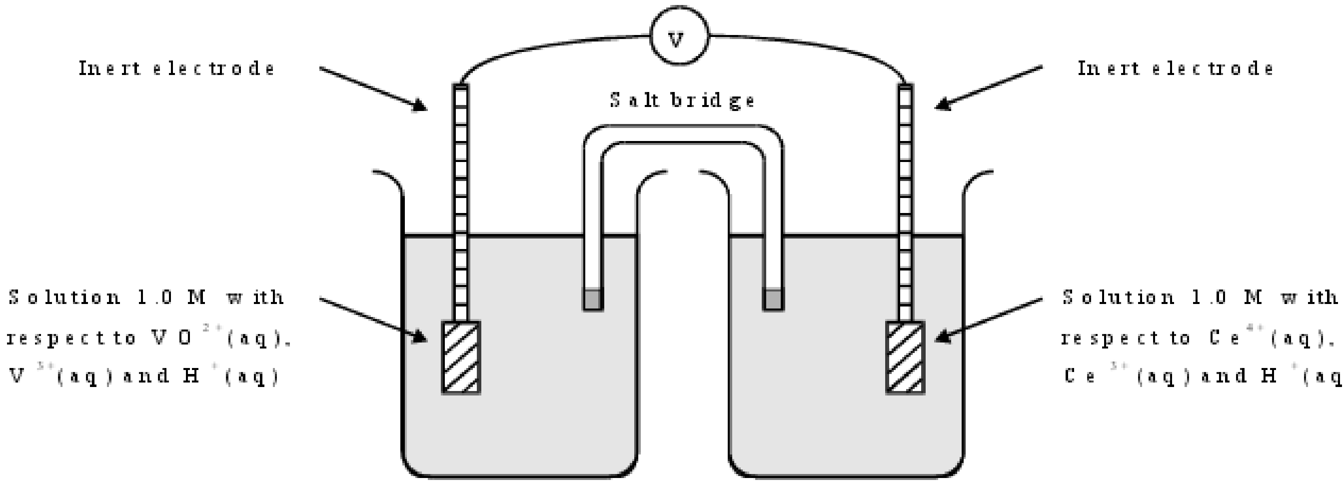
(b) A diagram of a cell is shown below.



(i) Calculate the overall standard potential of this cell.

(ii) State the polarity of the zinc electrode.

(c) A diagram of a cell is shown below.



(i) Calculate the overall standard potential of the cell.

(ii) Deduce the direction of electron flow in the external circuit when the inert electrodes are connected together.

(d) Using data from the table on page 6, derive an equation for the overall redox process which occurs when a solution containing Ce^{4+},(aq) is added to a solution containing V^{3+} (aq).

(e) Which, if any, of the four vanadium-containing species, V^{2+} (aq), V^{3+} (aq), VO^{2+} (aq) and VO^{2+} (aq), will convert Sn^{2+}(aq) into Sn^{4+} (aq) in acid solution?

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

						E°/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.

(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) .

(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.

4. The standard electrode potentials for some redox systems involving vanadium are shown below. These are labelled **A**, **B**, **C** and **D**.

		$E^\ominus$ / V
A	$VO_2^{+} + 2H^{+} + e^{-} \rightleftharpoons VO^{2+} + H_2O$	+1.00
B	$V^{3+} + e^{-} \rightleftharpoons V^{2+}$	–0.26
C	$V^{2+} + 2e^{-} \rightleftharpoons V$	–1.20
D	$VO^{2+} + 2H^{+} + e^{-} \rightleftharpoons V^{3+} + H_2O$	+0.34

(a) Which of the vanadium species shown in **A**, **B**, **C** and **D** is the most powerful oxidising agent?

(b) A student wishes to set up a cell with a standard cell potential of 0.60V.

(i) Which two of the redox systems, **A**, **B**, **C** or **D**, should he choose?

(ii) Complete the labelling of the following diagram which shows the cell with a standard cell potential of 0.60V.

(iii) The emf of this cell is only 0.60 V under standard conditions. What do you understand by the expression *standard conditions*?

PART 2.

1.(a)Define the term *oxidising agent* in terms of electrons.

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

Standard electrode potentials	$E^{\ominus}$/V
$2BrO_3^-(aq) + 12H^+(aq) + 10e^- \rightarrow Br_2(aq) + 6H_2O(l)$	+1.52
$Fe^{3+}(aq) + e^- \rightarrow F2^{2+}(aq)$	+0.77
$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$	+1.36
$O_3(g) + 2H^+(aq) + 2e^- \rightarrow O_2(g) + H_2O(l)$	+2.08
$F_2O(g) + 2H^+(aq) + 4e^- \rightarrow 2F^-(aq) + H_2O(l)$	+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).

(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.

3. Use the data below, where appropriate, to answer the following questions.

Standard electrode potentials	$E^{\ominus}$/V
$S_2O_8^{2-}(aq) + 2e^- \rightarrow 2SO_4^{2-}(aq)$	+2.01
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$	+1.51
$Cl_2(aq) + 2e^- \rightarrow 2Cl^-(aq)$	+1.36
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$	+1.33
$NO_3^-(aq) + 3H^+(aq) + 2e^- \rightarrow HNO_2(aq) + H_2O(l)$	+0.94
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.77

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

(b) Deduce the oxidation state of

(i) chromium in $Cr_2O_7^{2-}$

(ii) nitrogen in HNO_2

(c) (i) Calculate the e.m.f. of the cell represented by

$Pt | Mn^{2+}(aq), MnO_4^-(aq) || S_2O_8^{2-}(aq), SO_4^{2-}(aq) | Pt$

(ii) Deduce an equation for the reaction which occurs when an excess of $S_2O_8^{2-}$ (aq) is added to an aqueous solution of Mn^{2+}(aq) ions.

4. 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.

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

(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.

$Pt \mid H_2O_2(aq), O_2(g) \| IO_3^{-}(aq), I_2(aq) \mid Pt$

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

5. Use the table of standard electrode potentials given below to answer the following questions.

	$E^{\ominus}/V$
$Cl_2(g) + 2e^{-} \rightarrow 2Cl^{-}(aq)$	$+1.36$
$Br_2(l) + 2e^{-} \rightarrow 2Br^{-}(aq)$	$+1.07$
$NO_3^{-}(aq) + 3H^{+}(aq) + 2e^{-} \rightarrow HNO_2(aq) + H_2O(l)$	$+0.94$
$Fe^{3+}(aq) + e^{-} \rightarrow Fe^{2+}(aq)$	$+0.77$
$I_2(aq) + 2e^{-} \rightarrow 2I^{-}(aq)$	$+0.54$
$VO^{2+}(aq) + 2H^{+}(aq) + e^{-} \rightarrow V^{3+}(aq) + H_2O(l)$	$+0.34$
$V^{3+}(aq) + e^{-} \rightarrow V^{2+}(aq)$	-0.26
$Fe^{2+}(aq) + 2e^{-} \rightarrow Fe(s)$	-0.44

(a) In terms of electron transfer, define the term *oxidising agent.*

(b) (i) Give the conditions under which the electrode potential for $Cl_2(g)/2Cl^-(aq)$ is +1.36 V.

(ii) Give a change in one of these conditions which would result in the electrode potential becoming more positive. Explain your answer.

(c) (i) Which of the reducing agents in the table is the weakest?

(ii) Identify all the species in the table which could convert I^-aq) into I_2(aq) but which could not convert Br^-(aq) into Br_2(l).

(iii) Identify the metal ions which would be left in solution if an excess of powdered iron metal was added to an acidified solution containing VO^{2+}(aq) ions.

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

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

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

(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?

(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.

(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).

7. 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.

(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.

(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.

8. The table below shows some values for standard electrode potentials.

Electrode	Electrode reaction	$E^{\ominus}$ / V
A	$Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s)$	– 1.18
B	$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	– 0.44
C	$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	– 0.25
D	$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	– 0.14
E	$2H^+(g) + 2e^- \rightleftharpoons H_2(g)$	?

(a) (i) Give the name of electrode **E** and indicate its role in the determination of standard electrode potentials.

(ii) What is the value of the standard electrode potential for electrode **E**?

(b) The electrochemical cell set up between electrodes **C** and **D** can be represented by the cell diagram:

$$Ni(s)\,|\,Ni^{2+}(aq)\,||\,Sn^{2+}(aq)\,|\,Sn(s)$$

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

(ii) State which would be the positive electrode.

(iii) Write an equation to show the overall reaction in the cell.

(c) Use the standard electrode potential data given in the table above:

(i) to explain whether or not you would expect a reaction to occur if a piece of tin were to be added to a test tube containing aqueous iron(II) sulphate;

(ii) to predict and explain two observations you would expect to make if a small piece of manganese were to be added to a test tube containing hydrochloric acid of concentration 1 mol dm^{-3}.

9. Use the data below to answer the questions that follow

Reaction at 298 K			$E^{\ominus}$ */V*
$Ag^+(aq) + e^-$	$\rightarrow$	$Ag(s)+$	+0.08
$AgF(s) + e^-$	$\rightarrow$	$Ag(s) + F^-(aq)$	+0.78
$AgCl(s) + e^-$	$\rightarrow$	$Ag(s) + Cl^-(aq)$	+0.22
$AgBr(s) + e^-$	$\rightarrow$	$Ag(s) + Br^-(aq)$	+0.07
$H^+(aq) + e^-$	$\rightarrow$	$\frac{1}{2} H_2(g)$	0.00
$D^+ (aq) + e^-$	$\rightarrow$	$\frac{1}{2} D_2(g)$	–0.004
$AgI(s) + e^-$	$\rightarrow$	$Ag(s) + I^-(aq)$	–0.15

The symbol D denotes deuterium, which is heavy hydrogen, ${}^{2}_{1}H$.

(a) By considering electron transfer, state what is meant by the term *oxidising agent.*

(b) State which of the two ions, $H^+(aq)$ or $D^+(aq)$, is the more powerful oxidising agent. Write an equation for the spontaneous reaction which occurs when a mixture of aqueous H^+ ions and D^+ ions are in contact with a mixture of hydrogen and deuterium gas. Deduce the e.m.f.

of the cell in which this reaction would occur spontaneously.

(c) Write an equation for the spontaneous reaction which occurs when aqueous F^- ions and Cl^- ions are in contact with a mixture of solid AgF and solid AgCl. Deduce the e.m.f. of the cell in which this reaction would occur spontaneously.

(d) Silver does not usually react with dilute solutions of strong acids to liberate hydrogen.

(i) State why this is so.

(ii) Suggest a hydrogen halide which might react with silver to liberate hydrogen in aqueous solution. Write an equation for the reaction and deduce the e.m.f. of the cell in which this reaction would occur spontaneously

10. (a) The following reaction occurs in aqueous solution.

$$5S_2O_8^{2-} + Br_2 + 6H_2O \rightarrow 2BrO_3^- + 12H^+ + 10SO_4^{2-}$$

Identify the reducing agent in this reaction and write a half-equation for its action.

(b) The electrode potential for the half-equation

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

is measured by reference to a standard hydrogen electrode.

(i) State the temperature at which the standard electrode potential $E^\ominus$ is measured, and give the concentration of $Co^{2+}(aq)$ that must be used.

(ii) Electrode potentials are usually measured by reference to a secondary standard electrode. Identify a secondary standard electrode and give a reason why it is used rather than a standard hydrogen electrode.

(c) Cobalt in oxidation states +2 and +3 forms complex ions with water, ammonia and cyanide ligands. Use, where appropriate, the data given below to answer the questions which follow.

$[Co(H_2O)_6]^{3+}(aq) + e^- \rightarrow [Co(H_2O)_6]^{2+}(aq)$ $E^\ominus = +1.81$ V

$\frac{1}{2}O_2(g) + 2H^+(aq) + 2e^- \rightarrow H_2O(l)$ $E^\ominus = +1.23$ V

$[Co(NH_3)_6]^{3+}(aq) + e^- \rightarrow [Co(NH_3)_6]^{2+}(aq)$ $E^\ominus = +0.10$ V

$2H^+(aq) + 2e^- \rightarrow H_2(g)$ $E^\ominus = 0.00$ V

$[Co(CN)_6]^{3-}(aq) + e^- \rightarrow [Co(CN)_6]^{4-}(aq)$ $E^\ominus = -0.80$ V

(i) Which of the six cobalt species shown above is the most powerful oxidising agent?

(ii) Identify a cobalt(II) species which cannot be oxidised by gaseous oxygen.

(iii) Hydrogen is evolved when a salt containing the cobalt species $[Co(CN)_6]^{4-}(aq)$ is reacted with a dilute acid. Use the electrode potentials given above to explain the formation of the hydrogen gas.

11. The table below shows some values for standard electrode potentials. These data should be used, where appropriate, to answer the questions that follow concerning the chemistry of copper and iron.

Electrode reaction	**$E^\ominus$/V**
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	– 0.44
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+ 0.34
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$	+ 0.40
$NO_3^-(aq) + 4H^+(aq) + 3e^- \rightleftharpoons NO(g) + 2H_2O(l)$	+ 0.96

(a) Write an equation to show the reaction that occurs when iron is added to a solution of a copper(II) salt.

(b) A similar overall reaction to that shown in (a) would occur if an electrochemical cell was set up between copper and iron electrodes.

(i) Write down the cell diagram to represent the overall reaction in the cell.
(ii) Calculate the e.m.f. of the cell.

(c) (i) Use the standard electrode potential data given to explain why copper reacts with dilute nitric acid but has no reaction with dilute hydrochloric acid.
(ii) Write an equation for the reaction between copper and dilute nitric acid.

(d) Although iron is a widely used metal, it has a major disadvantage in that it readily corrodes in the presence of oxygen and water. The corrosion is an electrochemical process which occurs on the surface of the iron.
(i) Use the standard electrode potential data given to write an equation for the overall reaction that occurs in the electrochemical cell set up between iron, oxygen and water.
(ii) State, with a reason, whether the iron acts as the anode or cathode of the cell.
(iii) Predict and explain whether or not you would expect a similar corrosion reaction to occur with copper in the presence of oxygen and water.

12. (a) Name the standard reference electrode against which electrode potentials are measured and, for this electrode, state the conditions to which the term *standard* refers.

(b) The standard electrode potentials for two electrode reactions are given below.

$S_2O_8^{2-}(aq) + 2e^- \rightarrow 2SO_4^{2-}(aq)$ $\quad E^\ominus = +2.01\ V$

$Ag^+(aq) + e^- \rightarrow Ag(s)$ $\quad E^\ominus = +0.80\ V$

(i) A cell is produced when these two half-cells are connected.
Deduce the cell potential, $E^\ominus$, for this cell and write an equation for the spontaneous reaction.
(ii) State how, if at all, the electrode potential of the $S_2O_8^{2-}/SO_4^{2-}$ equilibrium would change if the concentration of SO_4^{2-} ions was increased.
Explain your answer.

13. For **each** of the reactions listed below
(i) identify which species, if any, are acting as oxidising agents;
(ii) determine the oxidation states before and after reaction of any species that are oxidised;
(iii) write half-equations, including state symbols, for all redox reactions that occur.

$2Cu^{2+}(aq) + 4I^-(aq) \rightarrow 2CuI(s) + I_2(aq)$

$5H_2O_2(aq) + 2Mn^{2+}(aq) \rightarrow 2MnO_4^-(aq) + 6H^+(aq) + 2H_2O(l)$

$Cr_2O_7^{2-}(aq) + H_2O(l) \rightarrow 2CrO_4^{2-}(aq) + 2H^+(aq)$

$Cl_2(aq) + 2OH^-(aq) \rightarrow Cl^-(aq) + ClO^-(aq) + H_2O(l)$