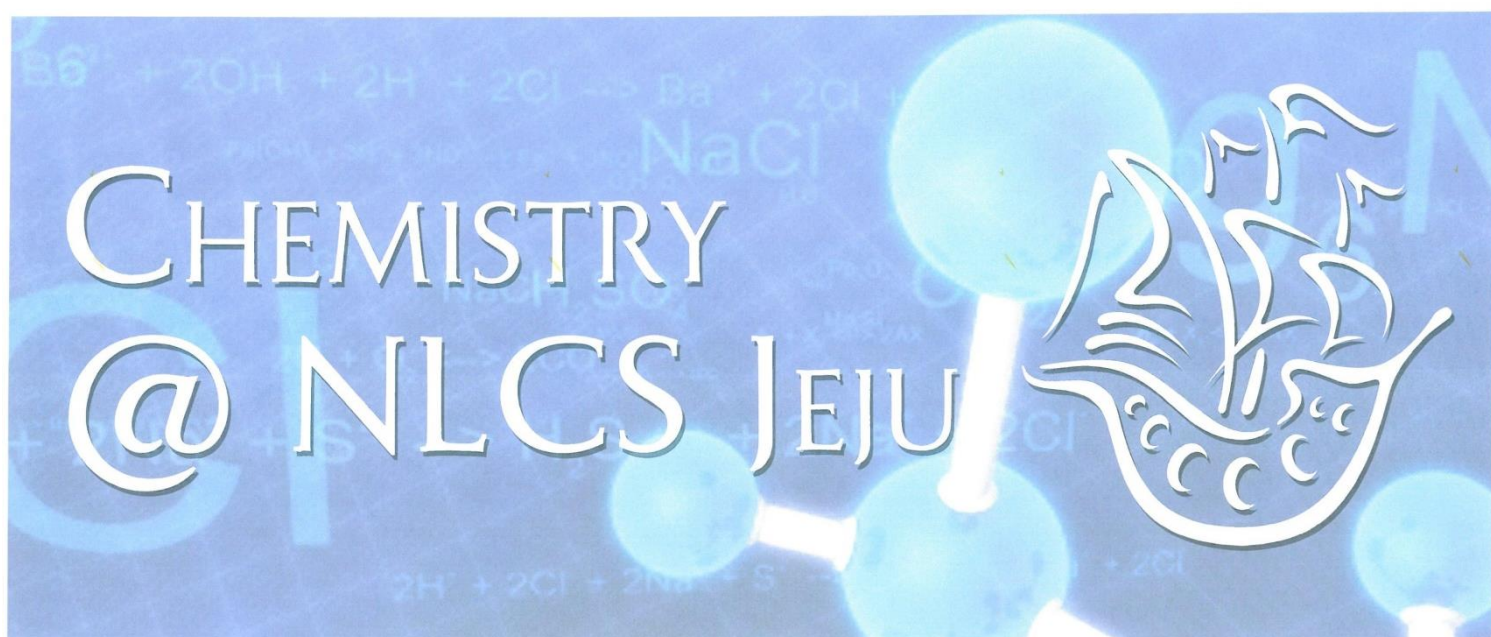


Topic 9

HIGHER Level



Summer & Winter Papers Summer 1999 to Summer 2013

Name: _____

Topic Exam Statistics (Paper 2):

Section	Marks	% of All Marks	Last four exams marks	Last four exams %
A	90/1080	8%	/160	%
B	353/2700	13%	/400	%
TOTAL	443/3780	12%	/560	%

Total number of papers represented here is 27, each with 40 marks of Section A and 100 marks of section B (4 questions from which you chose to answer only 2)

IB HL 9 EQ Paper 2 s99 to s13 incl W

HL SECTION A 10w

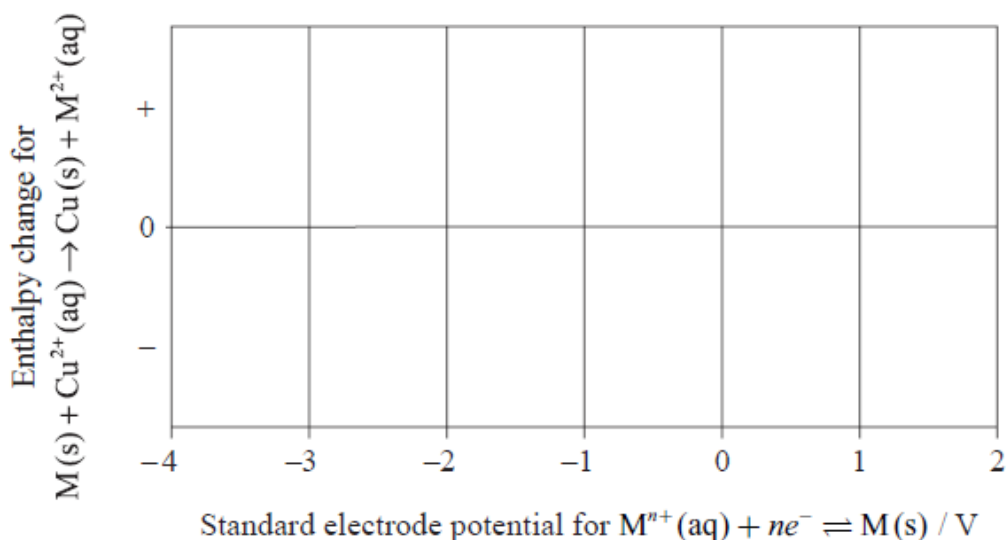
1 (Data Booklet)

(e) An experiment was designed to investigate how the enthalpy change for a displacement reaction relates to standard electrode potentials. The standard electrode potentials of some half reactions, $M^{n+}(aq) + ne^{-} \rightleftharpoons M(s)$, are listed in Table 14 of the Data Booklet. The following metals were available: copper, iron, magnesium, silver and zinc. Excess amounts of each metal were added to 1.00 mol dm^{-3} copper(II) sulfate solution. The temperature change was measured and the enthalpy change calculated.

(i) Suggest a possible hypothesis for the relationship between the enthalpy change that occurs when the metal, M, is added to copper(II) sulfate(aq) and the standard electrode potential for the half reaction $M^{n+}(aq) + ne^{-} \rightleftharpoons M(s)$. [1]

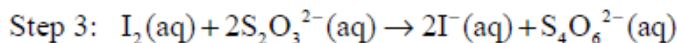
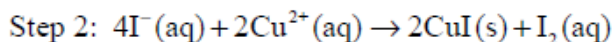
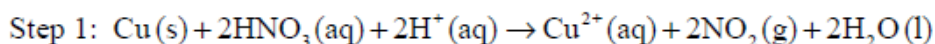
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(ii) Sketch a graph on the diagram below to illustrate your hypothesis. [2]



HL SECTION A 10s

1. Brass is a copper containing alloy with many uses. An analysis is carried out to determine the percentage of copper present in three identical samples of brass. The reactions involved in this analysis are shown below.



- (a) (i) Deduce the change in the oxidation numbers of copper and nitrogen in step 1. [2]

Copper:

.....

Nitrogen:

.....

- (ii) Identify the oxidizing agent in step 1. [1]

.....

- (b) A student carried out this experiment three times, with three identical small brass nails, and obtained the following results.

Mass of brass = 0.456 g ± 0.001 g

Titre	1	2	3
Initial volume of 0.100 mol dm ⁻³ S ₂ O ₃ ²⁻ (±0.05 cm ³)	0.00	0.00	0.00
Final volume of 0.100 mol dm ⁻³ S ₂ O ₃ ²⁻ (±0.05 cm ³)	28.50	28.60	28.40
Volume added of 0.100 mol dm ⁻³ S ₂ O ₃ ²⁻ (±0.10 cm ³)	28.50	28.60	28.40
Average volume added of 0.100 mol dm ⁻³ S ₂ O ₃ ²⁻ (±0.10 cm ³)	28.50		

(i) Calculate the average amount, in mol, of $S_2O_3^{2-}$ added in step 3. [2]

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(ii) Calculate the amount, in mol, of copper present in the brass. [1]

.....
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(iii) Calculate the mass of copper in the brass. [1]

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(iv) Calculate the percentage by mass of copper in the brass. [1]

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.....

(v) The manufacturers claim that the sample of brass contains 44.2 % copper by mass. Determine the percentage error in the result. [1]

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.....

(c) In step 1 the copper reacts to form a blue solution.

(i) State the full electronic configuration of Cu^{2+} . [1]

.....

(ii) Explain why the copper solution is coloured. [2]

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1. The data below is from an experiment used to determine the percentage of iron present in a sample of iron ore. This sample was dissolved in acid and all of the iron was converted to Fe^{2+} . The resulting solution was titrated with a standard solution of potassium manganate(VII), KMnO_4 . This procedure was carried out three times. In acidic solution, MnO_4^- reacts with Fe^{2+} ions to form Mn^{2+} and Fe^{3+} and the end point is indicated by a slight pink colour.

Titre	1	2	3
Initial burette reading / cm^3	1.00	23.60	10.00
Final burette reading / cm^3	24.60	46.10	32.50

Mass of iron ore / g	3.682×10^{-1}
Concentration of KMnO_4 solution / mol dm^{-3}	2.152×10^{-2}

- (a) Deduce the balanced redox equation for this reaction in **acidic** solution. [2]

.....

- (b) Identify the reducing agent in the reaction. [1]

.....

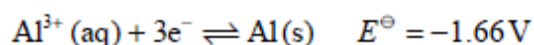
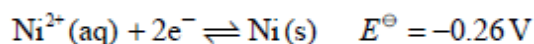
- (c) Calculate the amount, in moles, of MnO_4^- used in the titration. [2]

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4. (a) Outline two differences between an electrolytic cell and a voltaic cell. [2]

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(b) Consider the following half-cell reactions and their standard electrode potentials.



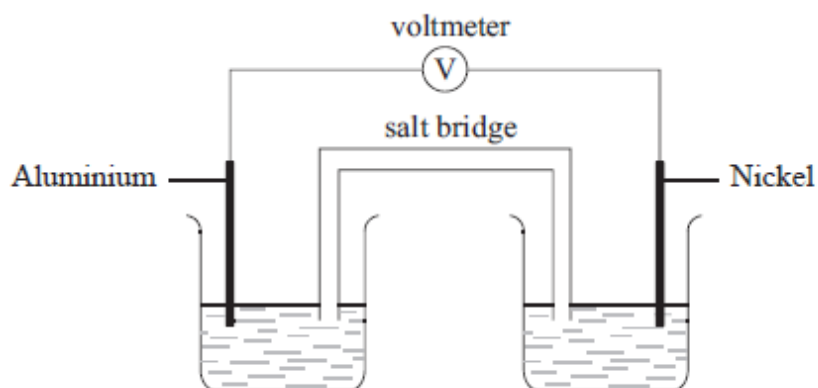
(i) Deduce a balanced equation for the overall reaction which will occur spontaneously when these two half-cells are connected. [2]

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.....

(ii) Determine the cell potential when the two half-cells are connected. [1]

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.....

(iii) On the cell diagram below, label the negative electrode (anode), the positive electrode (cathode) and the directions of the movement of electrons and ion flow. [4]



4. The percentage of iron(II) ions, Fe^{2+} , in a vitamin tablet can be estimated by dissolving the tablet in dilute sulfuric acid and titrating with standard potassium manganate(VII) solution, $\text{KMnO}_4(\text{aq})$. During the process iron(II) is oxidized to iron(III) and the manganate(VII) ion is reduced to the manganese(II) ion, $\text{Mn}^{2+}(\text{aq})$. It was found that one tablet with a mass of 1.43 g required 11.6 cm^3 of $2.00 \times 10^{-2} \text{ mol dm}^{-3} \text{ KMnO}_4(\text{aq})$ to reach the end-point.

(a) (i) State the half-equation for the oxidation of the iron(II) ions. [1]

.....

(ii) State the half-equation for the reduction of the MnO_4^- ions in acidic solution. [1]

.....

(iii) Deduce the overall redox equation for the reaction. [1]

.....

(b) (i) Calculate the amount, in moles, of MnO_4^- ions present in 11.6 cm^3 of $2.00 \times 10^{-2} \text{ mol dm}^{-3} \text{ KMnO}_4(\text{aq})$. [1]

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.....

(ii) Calculate the amount, in moles, of Fe^{2+} ions present in the vitamin tablet. [1]

.....

.....

(iii) Determine the percentage by mass of Fe^{2+} ions present in the vitamin tablet. [2]

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HL SECTION A 07w

4. (a) Iodide ions, $\text{I}^- (\text{aq})$, react with iodate ions, $\text{IO}_3^- (\text{aq})$, in an acidic solution to form molecular iodine and water.

(i) Determine the oxidation number of iodine in I^- and in IO_3^- . [1]

.....

(ii) Identify, with a reason, the species that undergoes: [2]

oxidation

.....

reduction

.....

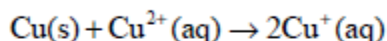
(iii) Write an ionic equation for the reaction of I^- with IO_3^- in an acidic solution. [2]

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(b) Use information from Table 15 of the Data Booklet to calculate the cell potential for the following reaction and state whether or not the reaction is spontaneous. [3]



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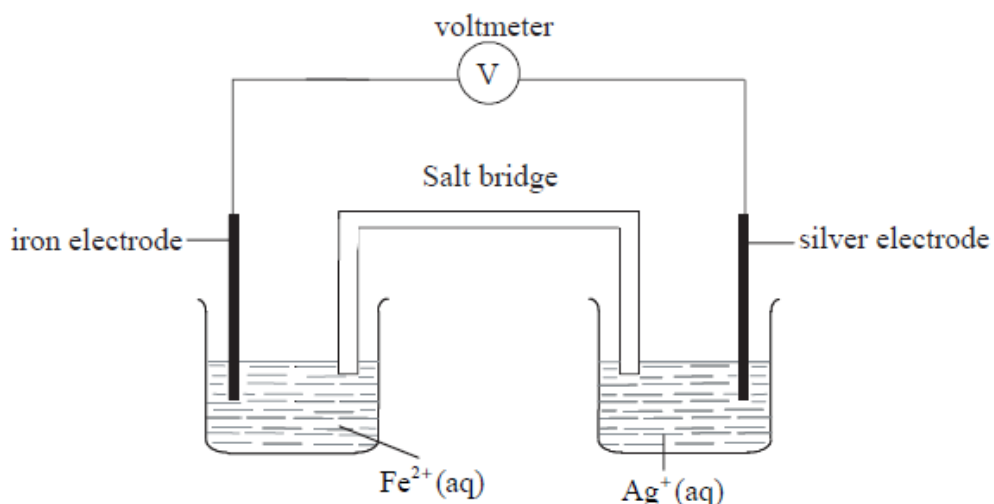
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HL SECTION A 07s

TABLE 15

2. The following diagram shows a voltaic cell.



(a) State an equation to represent the spontaneous reaction occurring in the cell. [1]

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(b) Define the term *standard electrode potential*. [1]

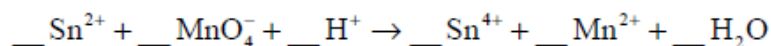
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(c) Use Table 15 from the Data Booklet to calculate the standard cell potential for the spontaneous reaction in (a). [1]

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(d) Draw arrows on the above diagram to indicate the direction of electron flow. [1]

3. Tin(II) ions can be oxidized to tin(IV) ions by acidified potassium permanganate(VII) solution according to the following unbalanced equation.



- (a) Identify the oxidizing agent and the reducing agent. [1]

Oxidizing agent.....

Reducing agent.....

- (b) Balance the equation above. [1]

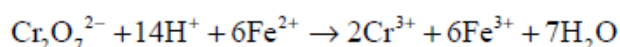
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HL SECTION A 06s

4. (a) (i) Define *oxidizing agent* in terms of electron transfer. [1]

.....

- (ii) Deduce the **change** in oxidation number of chromium in the reaction below. State with a reason whether the chromium has been oxidized or reduced. [2]



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- (b) Iron in food, in the form of Fe^{3+} , reacts with ascorbic acid (vitamin C), $\text{C}_6\text{H}_8\text{O}_6$, to form dehydroascorbic acid, $\text{C}_6\text{H}_6\text{O}_6$.

- (i) Write an ionic half-equation to show the conversion of ascorbic acid to dehydroascorbic acid in aqueous solution. [1]

.....

- (ii) In the other ionic half-equation Fe^{3+} is converted to Fe^{2+} . Deduce the overall equation for the reaction between $\text{C}_6\text{H}_8\text{O}_6$ and Fe^{3+} . [1]

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5. (a) When a concentrated aqueous solution of sodium chloride is electrolyzed using inert electrodes, a different gas is produced at each electrode.

(i) Write equations for the oxidation and reduction half-reactions. [2]

Oxidation half-reaction:

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Reduction half-reaction:

.....

(ii) Explain why sodium is not formed during the electrolysis of aqueous NaCl solution. [1]

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(b) Deduce the products formed during the electrolysis of an aqueous solution of sodium fluoride. Write an equation for the reaction at the positive electrode (the anode) and give your reasoning. [4]

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4. A half-cell (A) is set up by placing a platinum electrode in a solution containing both Fe^{2+} and Fe^{3+} ions at a concentration of 1 mol dm^{-3} . This half-cell is then connected by means of a salt bridge to another half-cell (B) containing an iron electrode in a 1 mol dm^{-3} solution of Fe^{2+} ions.

(a) State the function of the salt bridge. [1]

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(b) The two electrodes are connected externally.

(i) Use Table 15 of the Data Booklet to determine the cell standard electrode potential. [2]

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.....
.....

(ii) Give the redox reactions that occur in each half-cell. [2]

A:

B:

(iii) State the direction of the electron flow in the external circuit. [1]

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HL SECTION A 01w

4. A blue aqueous solution of copper(II) sulfate, CuSO_4 , is electrolysed using copper electrodes.

(a) Write balanced half-equations, with state symbols, for the reactions occurring at the:

anode (positive electrode): [1]

cathode (negative electrode): [1]

(b) State whether or not the colour of the solution will change as the electrolysis proceeds. Explain your answer. [1]

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.....

(c) Write a balanced chemical equation for the products formed if the copper anode is replaced by a graphite anode. [2]

.....

(d) State whether or not the colour intensity and pH of the solution will change as the electrolysis proceeds in (c) above. [2]

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(e) Calculate the mass of copper produced when a current of 0.180 amperes is passed through a 1.0 mol dm^{-3} copper sulfate solution for 20 minutes 10 seconds. [3]

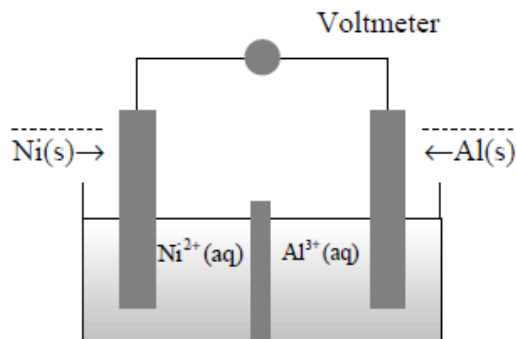
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HL SECTION A 00s

4. (a) Define a reducing agent in terms of electrons. [1]

.....

(b) In the following representation of a cell, label each electrode with a + or a - sign, as appropriate, and draw an arrow on the connecting wire to indicate the direction of electron flow. (Refer to Table 15 of the Chemistry Data Booklet.) [2]



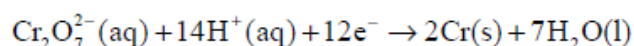
(c) (i) Write the balanced equation for the spontaneous reaction in the above cell. [2]

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(ii) Calculate the standard cell potential. [2]

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(d) Chromium is deposited from an acidic solution containing the dichromate(VI) ion, according to the equation

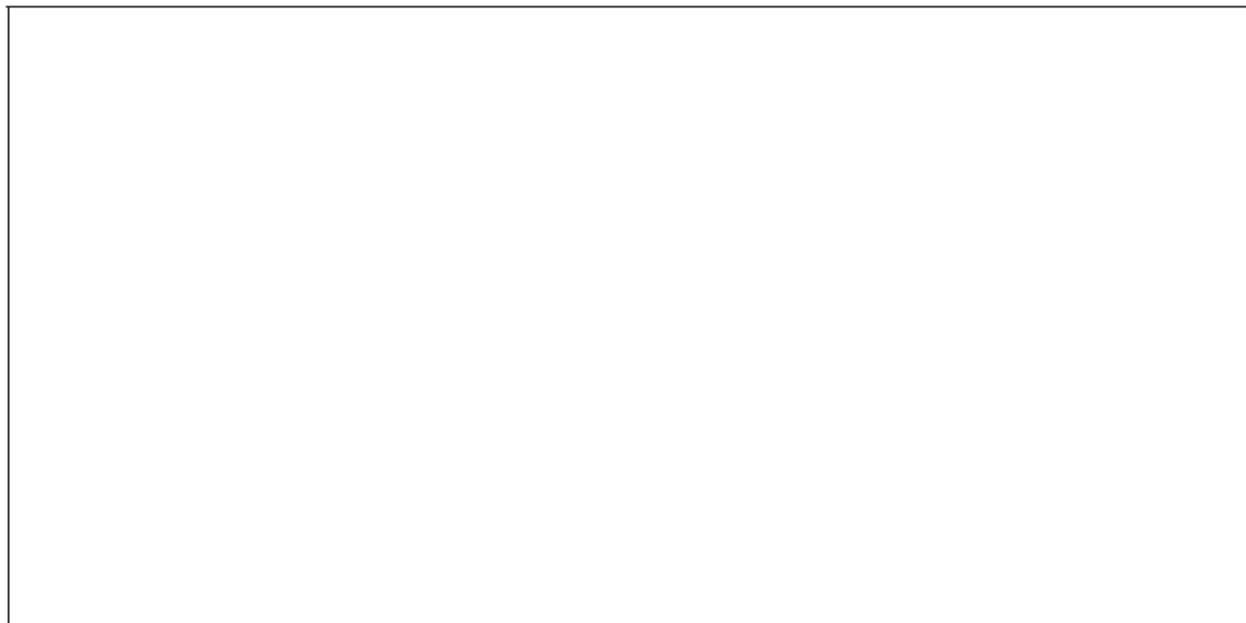


How many moles of Cr(s) can be deposited by a current of 8.00 A flowing for 2.00 hours? [5]

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(c) Freshly prepared iron(II) bromide can be electrolysed both in the liquid state and in aqueous solution.

(i) Describe, using a diagram, the essential components of an electrolytic cell. [3]



(ii) Describe the **two** ways in which current is conducted in an electrolytic cell. [2]

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(iii) Predict and explain the products of electrolysis of a **dilute** iron(II) bromide solution. [4]

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(iv) Identify another product that is formed if the solution of iron(II) bromide is **concentrated**. [1]

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(v) Explain why this other product is formed. [1]

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4. Arsenic and nitrogen play a significant role in environmental chemistry. Arsenous acid, H_3AsO_3 , can be found in oxygen-poor (anaerobic) water, and nitrogen-containing fertilizers can contaminate water.

(a) (i) Define *oxidation* and *reduction* in terms of electron loss or gain. [1]

Oxidation:
Reduction:

(ii) Deduce the oxidation numbers of arsenic and nitrogen in each of the following species. [4]

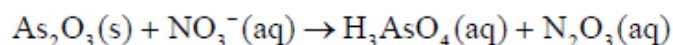
As_2O_3 :
NO_3^- :
H_3AsO_3 :
N_2O_3 :

(iii) Distinguish between the terms *oxidizing agent* and *reducing agent*. [1]

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- (iv) In the removal of arsenic from contaminated groundwater, H_3AsO_3 is often first oxidized to arsenic acid, H_3AsO_4 .

The following **unbalanced** redox reaction shows another method of forming H_3AsO_4 .



Deduce the balanced redox equation in **acid**, and then identify both the oxidizing and reducing agents. [3]

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- (b) The electrolysis of aqueous copper(II) sulfate is an example of an electrolysis process where the nature of the electrodes can determine which products form. Platinum electrodes were used in **process 1** and copper electrodes in **process 2**.

- (i) Draw an annotated diagram of the electrolytic cell in **process 1** and identify the direction of electron flow. [2]

- (ii) For **process 1** (platinum electrodes), state the half-equations occurring at the positive electrode (anode) and negative electrode (cathode). **Include state symbols for all species.** Describe what is observed at each electrode and comment on any change in the colour and the acidity of the solution.

[7]

Half-equation at positive electrode (anode):

.....
.....

Half-equation at negative electrode (cathode):

.....
.....

Observation at positive electrode (anode):

.....
.....

Observation at negative electrode (cathode):

.....
.....

Change in colour (if any) of the solution:

.....
.....

Change in acidity (if any) of the solution:

.....
.....

- (iii) For **process 2** (copper electrodes), state the half-equations occurring at the positive electrode (anode) and negative electrode (cathode). **Include state symbols for all species.** Describe what is observed at each electrode and comment on any change in the colour and the acidity of the solution.

[7]

Half-equation at positive electrode (anode):

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Half-equation at negative electrode (cathode):

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Observation at positive electrode (anode):

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Observation at negative electrode (cathode):

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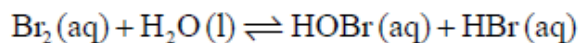
Change in colour (if any) of the solution:

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Change in acidity (if any) of the solution:

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- (e) When bromine dissolves in water, 1 % of the original bromine molecules react according to the following equation.



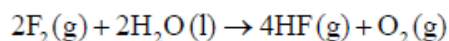
- (i) Deduce the oxidation numbers of bromine in the reactant **and** products. [2]

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- (ii) Explain the changes in the oxidation numbers of bromine. [1]

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- (f) Fluorine reacts with water to produce oxygen.



- (i) Identify the oxidizing agent in the reaction. [1]

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- (ii) 100 cm³ of fluorine gas is added to water. Calculate the volume of oxygen produced at the same temperature and pressure. [1]

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7. Chromium is a typical transition metal with many uses.

(a) Distinguish between the terms *oxidation* and *reduction* in terms of oxidation numbers. [1]

.....
.....

(b) State the names of Cr_2O_3 and CrO_3 . [2]

Cr_2O_3 :

CrO_3 :

(c) (i) Define the term *oxidizing agent*. [1]

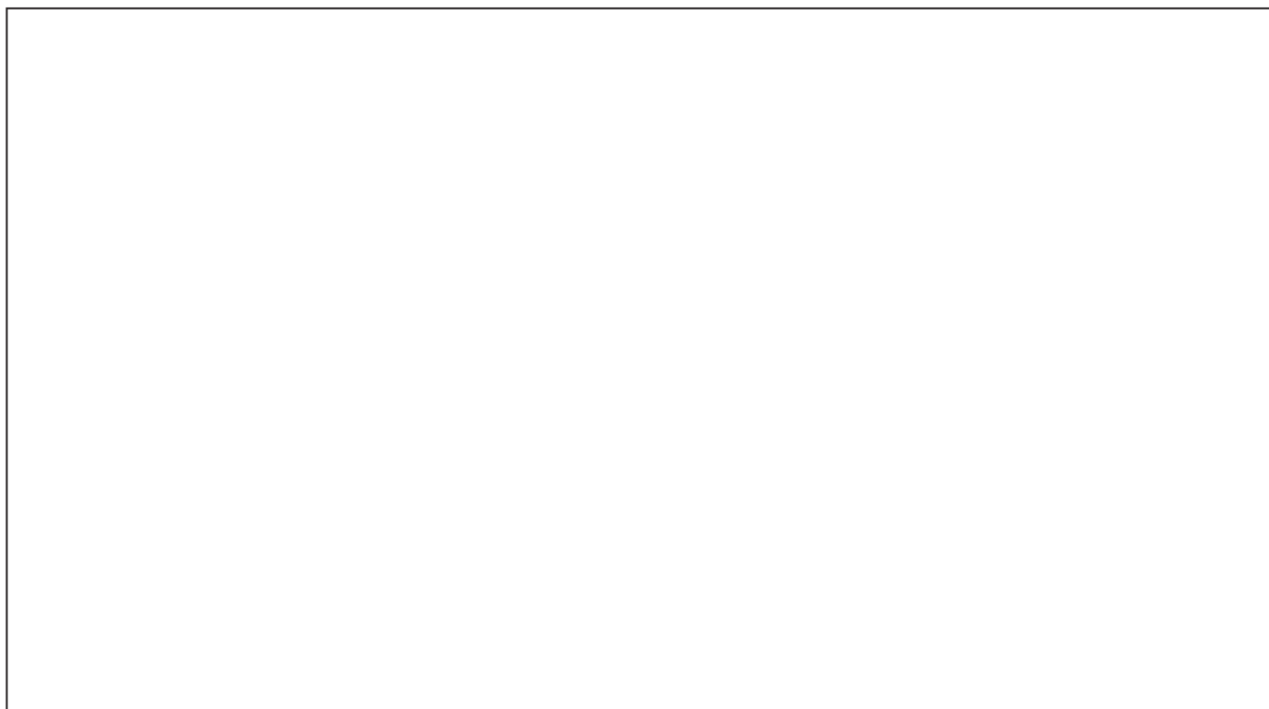
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(ii) $\text{Cr}_2\text{O}_7^{2-}(\text{aq})$ and $\text{I}^-(\text{aq})$ ions react together in the **presence of acid** to form $\text{Cr}^{3+}(\text{aq})$ and $\text{IO}_3^-(\text{aq})$ ions. Deduce the balanced chemical equation for this redox reaction and identify the species that acts as the oxidizing agent. [3]

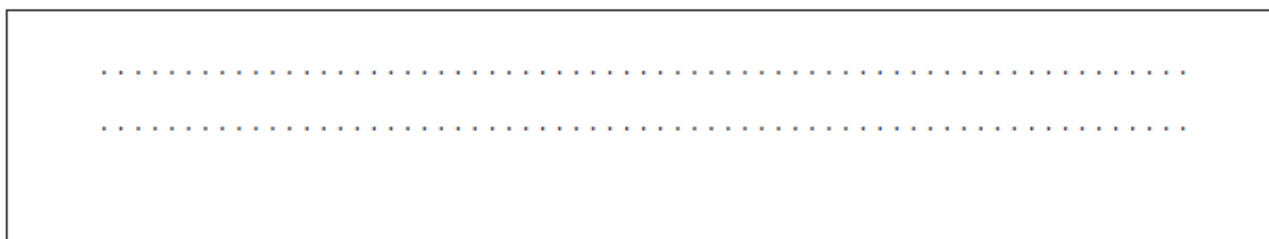
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(d) A voltaic cell is constructed as follows. One half-cell contains a platinum electrode in a solution containing $\text{K}_2\text{Cr}_2\text{O}_7$ and H_2SO_4 . The other half-cell contains an iron electrode in a solution containing Fe^{2+} ions. The two electrodes are connected to a voltmeter and the two solutions by a salt bridge.

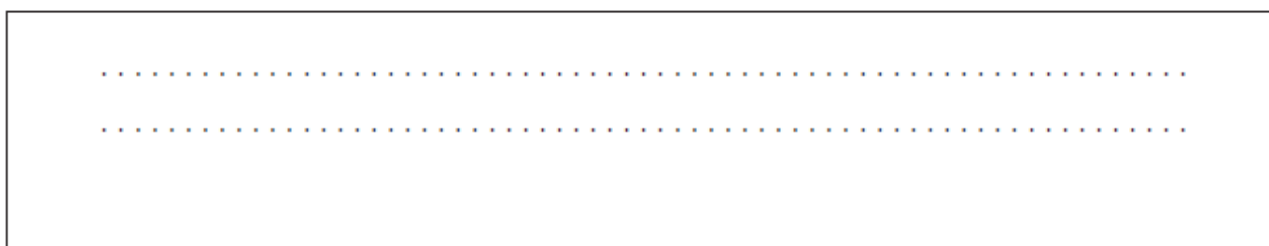
(i) Draw a diagram of the voltaic cell, labelling the positive and negative electrodes (cathode and anode) and showing the direction of movement of the electrons and ions. Deduce an equation for the reaction occurring in each of the half-cells, and the equation for the overall cell reaction. [5]



(ii) Define the term *standard electrode potential*. [1]



(iii) Calculate the cell potential, in V, under standard conditions, using information from Table 14 of the Data Booklet. [1]



(f) Chromium is often used in electroplating. State what is used as the positive electrode (anode), the negative electrode (cathode) and the electrolyte in the chromium electroplating process. [3]

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HL B 11s

TABLE 14 of Data Booklet is needed

9. (a) The conditions used in an electrolytic cell can determine the products formed.

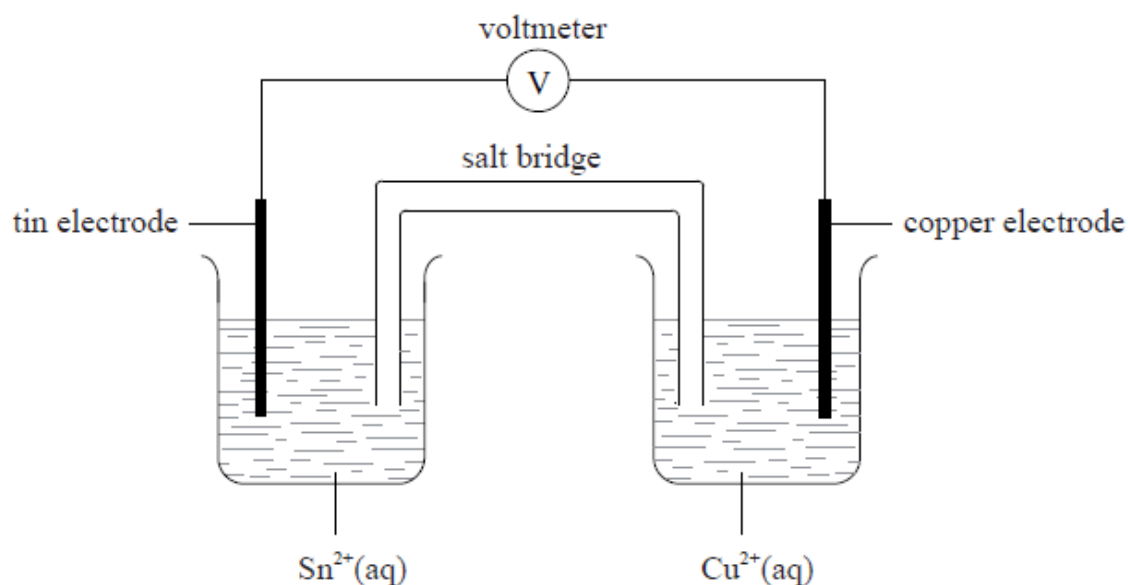
- (i) Draw an electrolytic cell illustrating the electrolysis of molten nickel(II) bromide, NiBr_2 . Include in the diagram the direction of the electron flow, the polarity of electrodes and state the half-equations for the product formed at each electrode. [5]

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.....

- (ii) Deduce the equations for the formation of the major product at the positive electrode (anode) when the following aqueous solutions are electrolysed.
- dilute sodium chloride
 - concentrated sodium chloride
- [2]

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(b) A voltaic cell is constructed from two half-cells as illustrated below.



(i) Use Table 14 of the Data Booklet to deduce the equation for the spontaneous reaction occurring in this cell. [1]

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.....

(ii) Calculate the standard potential for this cell. [1]

.....
.....

(iii) State the conditions necessary for the potential of the cell to equal that calculated in part (b) (ii) using the data from Table 14. [1]

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.....

- (c) Using the data below and data from Table 14 of the Data Booklet, predict and explain which metal, cadmium or chromium, may be obtained by electrolysis of separate aqueous solutions of $\text{Cd}^{2+}(\text{aq})$ ions and $\text{Cr}^{2+}(\text{aq})$ ions. [2]

	E^\ominus / V
$\text{Cd}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cd}(\text{s})$	-0.40
$\text{Cr}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cr}(\text{s})$	-0.91

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- (d) (i) Electrolysis is used in the electroplating of metals. The same amount of current is passed through separate aqueous solutions of NiSO_4 , $\text{Sn}(\text{SO}_4)_2$ and $\text{Cr}_2(\text{SO}_4)_3$ in separate electrolytic cells for the same amount of time. State and explain which cell would deposit the greatest amount (in mol) of metal. Identify the electrode at which the metal is deposited. [3]

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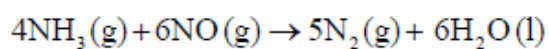
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- (ii) For the $\text{Sn}(\text{SO}_4)_2$ cell, suggest **two** factors, other than time and current, that would affect the amount of metal deposited during electroplating. [2]

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- (e) Nitrogen monoxide may be removed from industrial emissions via a reaction with ammonia as shown by the equation below.



- (i) Deduce the oxidation number of the nitrogen in the reactants and product. [3]

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- (ii) Deduce the oxidation and reduction half-equations and identify the oxidizing agent for the reaction. [3]

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HL B 10w

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- (b) Magnesium can be produced from the electrolysis of molten magnesium chloride, MgCl_2 .

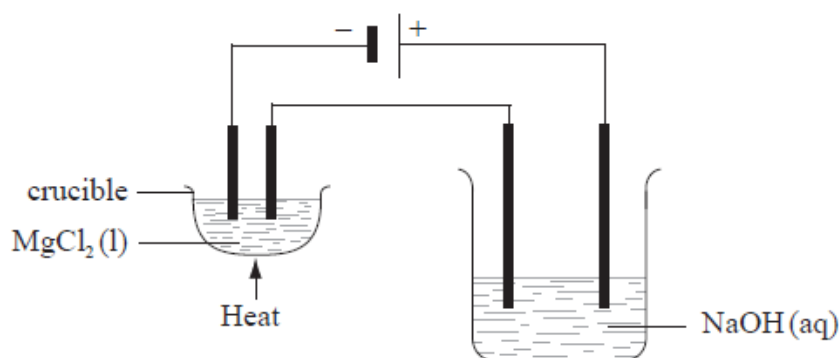
- (i) Explain how molten magnesium chloride conducts an electric current. [2]
- (ii) Identify the electrode where oxidation occurs during electrolysis of molten magnesium chloride and state an equation for the half-reaction. [2]
- (iii) Explain why magnesium is not formed during the electrolysis of aqueous magnesium chloride solution. [1]

HL B 10s

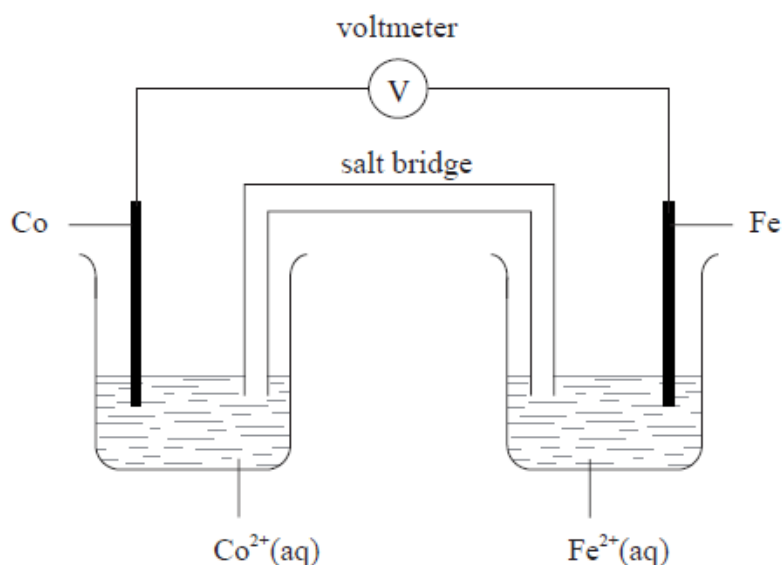
6. (a) (i) Draw an annotated diagram of a voltaic cell composed of a magnesium electrode in 1.0 mol dm^{-3} magnesium nitrate solution and a silver electrode in 1.0 mol dm^{-3} silver nitrate solution. State the direction of electron flow on your diagram. [4]
- (ii) Deduce half-equations for the oxidation and reduction reactions. [2]
- (iii) Using Table 14 of the Data Booklet, calculate the cell potential for this cell. [2]
- (b) The standard electrode potentials for three other electrode systems are given below.

	E^\ominus / V
$\text{MnO}_4^- (\text{aq}) + 8\text{H}^+ (\text{aq}) + 5\text{e}^- \rightleftharpoons \text{Mn}^{2+} (\text{aq}) + 4\text{H}_2\text{O} (\text{l})$	+1.51
$\text{Fe}^{3+} (\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+} (\text{aq})$	+0.77
$\text{Cd}^{2+} (\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cd} (\text{s})$	-0.40

- (i) Identify which species in the table above is the best reducing agent. [1]
- (ii) Deduce the equation for the overall reaction with the greatest cell potential. [2]
- (c) These values were obtained using a standard hydrogen electrode. Describe the materials and conditions used in the standard hydrogen electrode. (A suitably labelled diagram is acceptable). [4]
- (d) (i) Solid sodium chloride does not conduct electricity but molten sodium chloride does. Explain this difference. [2]
- (ii) Outline what happens in an electrolytic cell during the electrolysis of molten sodium chloride using inert electrodes. Deduce equations for the reactions occurring at each electrode. [4]
- (iii) Two electrolytic cells are connected in series as shown in the diagram below. In one there is molten magnesium chloride and in the other, dilute sodium hydroxide solution. Both cells have inert electrodes. If 12.16 g of magnesium is produced in the first cell, deduce the identity and mass of products produced at the positive and negative electrodes in the second cell. [4]



7. (a) An electrochemical cell is made from an iron half-cell connected to a cobalt half-cell:



The standard electrode potential for $\text{Fe}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Fe}(\text{s})$ is -0.45 V . The total cell potential obtained when the cell is operating under standard conditions is 0.17 V . Cobalt is produced during the spontaneous reaction.

- (i) Define the term *standard electrode potential* and state the meaning of the minus sign in the value of -0.45 V . [3]
 - (ii) Calculate the value for the standard electrode potential for the cobalt half-cell. [1]
 - (iii) Deduce which species acts as the oxidizing agent when the cell is operating. [1]
 - (iv) Deduce the equation for the spontaneous reaction taking place when the iron half-cell is connected instead to an aluminium half-cell. [2]
 - (v) Explain the function of the salt bridge in an electrochemical cell. [2]
- (b) Deduce the oxidation number of cobalt in the following species.
- (i) $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ [1]
 - (ii) $\text{Co}_2(\text{SO}_4)_3$ [1]
 - (iii) $[\text{CoCl}_4]^{2-}$ [1]

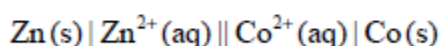
- (c) An electrolytic cell is made using a very dilute solution of sodium chloride.
- Draw a labelled diagram of the cell. Use an arrow to show the direction of the electron flow and identify the positive and negative electrodes. [3]
 - Give the formulas of all the ions present in the solution. [2]
 - Predict the products obtained at each electrode and state the half-equation for the formation of each product. [3]
 - Deduce the molar ratios of the products obtained at the two electrodes. [1]
- (d) Predict the products by giving the relevant half-equation for the reaction occurring at each electrode if the electrolyte of the cell described in part (c) was changed to:
- concentrated sodium chloride [2]
 - molten sodium bromide [2]

HL B 08w

DATA BOOKLET

6

- (e) Some standard electrode potentials are shown in Table 15 of the Data Booklet.
- From this table, identify a species that will reduce bromine to bromide ions but not iodine to iodide ions under standard conditions. Deduce the redox equation for the spontaneous reaction that occurs. [2]
 - Calculate the cell potential of a cell set up by connecting half-cells of aluminium and silver together under standard conditions. [1]
 - The cell potential of the cell represented below, under standard conditions is +0.48 V.



Deduce the standard electrode potential for the following half-reaction.

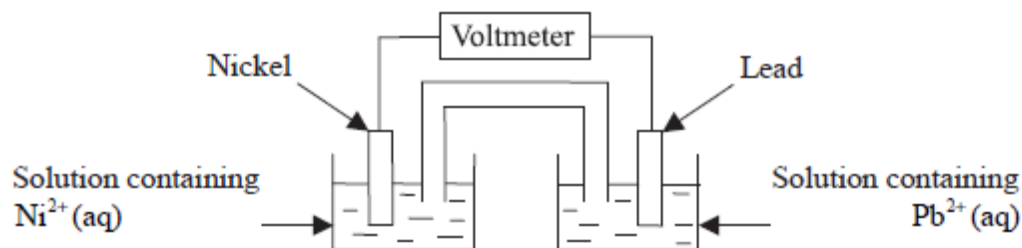


HL B 08s

8. (a) Electrolysis of an aqueous solution of copper(II) sulfate, CuSO_4 , can be carried out using platinum electrodes.
- State an equation for the half-reaction occurring at the positive electrode (anode) and **one** observation that could be made as a result. [2]
 - State an equation for the half-reaction occurring at the negative electrode (cathode) and **one** observation that could be made as a result. [2]
 - Describe **two** changes or observations in the electrolyte as result of these half reactions. [2]
 - Determine the relative amount, in moles, of products formed at each electrode. [1]
 - Identify another compound which will form the same products at the positive and negative electrodes. [1]
 - The same process is carried out using copper electrodes instead of the platinum electrodes. Describe the changes or observations that take place at both the electrodes and in the electrolyte. [3]
- (b) Identify **two** factors that affect the quantity of copper produced during the electrolysis of an aqueous copper(II) sulfate solution. [1]
- (c) Consider the following half cell reactions and their standard electrode potentials.
- $$\text{Ni}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Ni}(\text{s}) \quad E^{\ominus} = -0.23\text{V}$$
- $$\text{MnO}_4^{-}(\text{aq}) + 8\text{H}^{+}(\text{aq}) + 5\text{e}^{-} \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l}) \quad E^{\ominus} = +1.51\text{V}$$
- State the conditions needed for the electrode potentials to be described as standard. [1]
 - Deduce a balanced equation for the overall reaction which will occur spontaneously when the two half cells are connected. [2]
 - Identify the reducing agent in the above reaction and determine the change in oxidation number for the **oxidizing agent**. [2]
 - Determine the cell potential when the two half cells are connected. [1]
 - Draw and label a diagram of the voltaic cell from part (c). Indicate the anode, cathode, the direction of the electron movement and ion flow. [5]
- (d) Outline **two** differences between an electrolytic cell and a voltaic cell. [2]

(c) Use information from Table 15 of the Data Booklet, where relevant, in answering this part.

(i) The diagram shows the apparatus used in an experiment involving half-cells.



The reading on the voltmeter is 0.10 V. State two standard conditions that apply for this reading to be obtained.

[2]

(ii) The voltmeter is replaced by a 2 volt power supply, so that non-spontaneous reactions occur. Deduce the half-equations for these reactions.

[2]

(iii) Chlorine gas is formed when potassium manganate(VII) is added to concentrated hydrochloric acid. Calculate the cell potential for this reaction and deduce the equation for the reaction.

[3]

(iv) Explain why potassium dichromate(VI) does not react with concentrated hydrochloric acid.

[1]

HL B 05w

TABLE 15

7. (a) Electrolysis can be used to obtain chlorine from molten sodium chloride. Write an equation for the reaction occurring at each electrode and describe the two different ways in which electricity is conducted when the cell is in operation. [4]
- (b) In one experiment involving the electrolysis of molten sodium chloride, 0.1 mol of chlorine was formed. Deduce, giving a reason, the amount of sodium formed at the same time. [2]
- (c) In another experiment involving the electrolysis of molten sodium chloride, the time of the electrolysis was halved and the current increased from 1 amp to 5 amp, compared to the experiment in (b). Deduce the amount of chlorine formed, showing your working. [2]
- (d) If dilute aqueous sodium chloride is electrolyzed, a different product is obtained at each electrode. Identify the product formed at each electrode and write an equation showing its formation. [4]
- (e) Two reactions occurring in the manufacture of bromine from sea water are
- I $\text{Cl}_2(\text{g}) + 2\text{Br}^-(\text{aq}) \rightarrow 2\text{Cl}^-(\text{aq}) + \text{Br}_2(\text{g})$
- II $\text{Br}_2(\text{g}) + \text{SO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{HBr}(\text{g}) + \text{H}_2\text{SO}_4(\text{g})$
- (i) Explain, by reference to electrons, why reaction I is referred to as a redox reaction. [2]
- (ii) State and explain whether SO_2 is reduced or oxidized in reaction II by referring to the oxidation numbers of sulfur in this reaction. [3]
- (f) Define the term *standard electrode potential* of an element. Table 15 of the Data Booklet contains E^\ominus values for two reactions involving $\text{O}_2(\text{g})$. Identify the E^\ominus value of the reaction that could be used to oxidize bromide ions and explain your reasoning. Construct a balanced equation for the oxidation of bromide ions using this reaction and calculate the cell potential. [8]

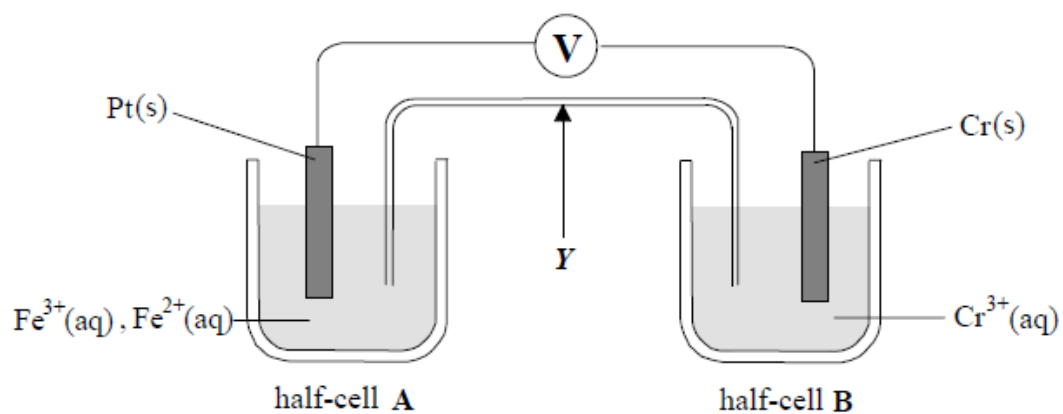
HL B 04w

9. The following are standard electrode potentials.

Half-equation	E^\ominus / V
$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Zn}(\text{s})$	-0.76
$\text{Cr}^{3+}(\text{aq}) + 3\text{e}^- \rightleftharpoons \text{Cr}(\text{s})$	-0.74
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Fe}(\text{s})$	-0.44
$\text{Sn}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Sn}(\text{s})$	-0.14
$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu}(\text{s})$	+0.34
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	+0.77

- (a) These values were obtained using a standard hydrogen electrode. Describe the materials and conditions used in the standard hydrogen electrode. (A suitably labelled diagram is acceptable.) [5]
- (b) Define the term *oxidizing agent* in terms of electron transfer and identify the strongest oxidizing agent in the list above. [2]
- (c) A cell was set up using zinc in zinc sulfate solution and copper in copper(II) sulfate solution, both solutions being under standard conditions.
- (i) Calculate the cell potential. [1]
- (ii) Write an equation for the spontaneous cell reaction. [2]
- (d) Both zinc and tin are used to coat iron to prevent it from rusting. Once the surface is scratched, oxygen and water containing dissolved ions come into contact with the iron and the coating metal.
- (i) State and explain whether zinc or tin would be more effective in preventing iron from rusting under these conditions. [2]
- (ii) Electroplating may be used to coat one metal with another metal. Identify the **three** factors affecting the amount of metal discharged during electroplating. [3]
- (iii) Explain why electrolysis of aqueous zinc sulfate is not used for coating with zinc metal. [2]

(e) Another cell was set up as shown below.



- (i) Identify the part of the cell labelled *Y* and outline its function. [2]
- (ii) Write an equation for the initial reactions at each electrode and hence write an equation for the cell reaction. [4]
- (iii) Describe the direction of electron flow in the external circuit. [1]
- (iv) Calculate the cell potential. [1]

HL B 04s

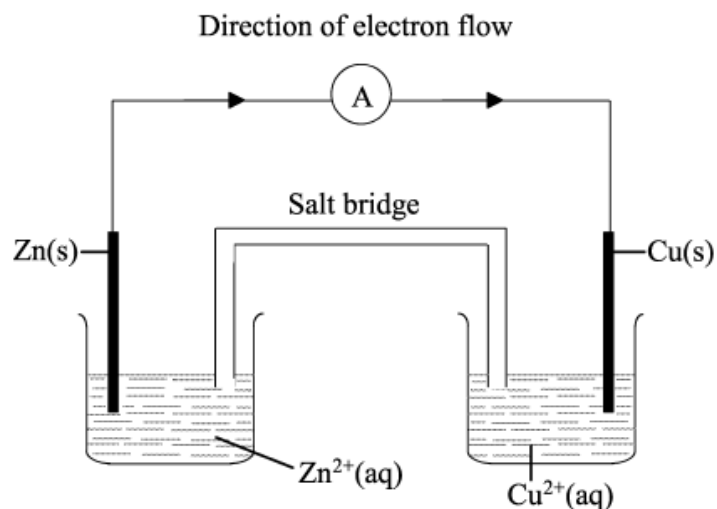
TABLE 15

10. (a) Some standard electrode potentials are shown in Table 15 of the Data Booklet.
- (i) State three conditions under which the hydrogen electrode is assigned a potential of zero. [3]
 - (ii) Calculate the cell potential of a cell made by connecting standard copper and zinc electrodes. State the direction of electron flow in the external circuit when the cell produces current. Outline the changes occurring at the electrodes and in the solutions during the process. [5]
- (b) Using information from Table 15, determine whether or not there is a spontaneous reaction between copper metal and a solution containing hydrogen ions. [2]
- (c) Using information from Table 15, identify a substance that will oxidize bromide ions but not chloride ions. Explain your choice, and write an equation for the redox reaction you have chosen. [5]
- (d) A current is passed through molten sodium chloride. Identify the substance formed at each electrode and write an equation to represent the formation of each substance. Determine the mole ratio in which the substances are formed. [5]
- (e) Sodium chloride in aqueous solution is electrolysed.
- (i) Identify the substances formed, and their relative amounts, when a concentrated solution is used. [2]
 - (ii) Identify the substances formed, and their relative amounts, when a very dilute solution is used. [2]
 - (iii) Write an equation for a reaction occurring when aqueous sodium chloride, but not molten sodium chloride, is electrolysed. [1]

HL B 03s

DATA BOOKLET NEEDED

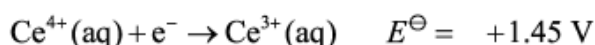
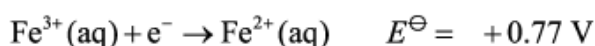
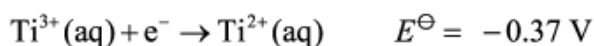
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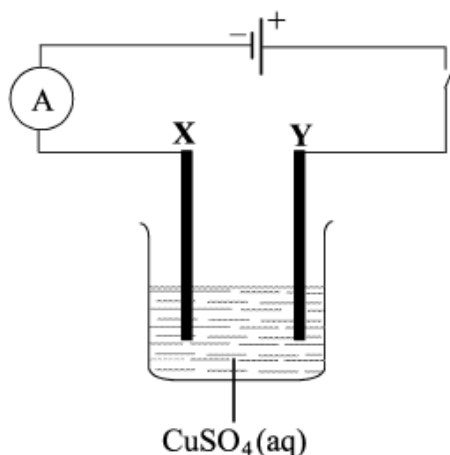
(a) The apparatus shown above may be used to carry out a redox reaction.

- (i) State the function of the salt bridge. [1]
- (ii) Write a half-equation for the oxidation reaction. [1]
- (iii) The above reactions are carried out under *standard conditions*. State what the standard conditions are for the cell. [2]
- (iv) Using the Data Booklet, calculate the cell potential for the above cell. [2]
- (v) State and explain what happens to the concentration of the copper(II) ions when the cell is producing an electric current. [2]
- (vi) State **two** observations that could be made if the zinc rod were placed in a solution of copper(II) ions. [2]

(b) The standard electrode potentials for three electrode systems are given below.



- (i) Using the data above, deduce which species is the best reducing agent, giving a reason in terms of electrons for your answer. [2]
- (ii) Write an equation, including state symbols, for the overall reaction with the greatest cell potential. [2]
- (iii) State and explain the sign of ΔG^{\ominus} for the reaction in (b) (ii). [2]
- (c) (i) State the name of a solution that would produce **only** hydrogen and oxygen when electrolyzed using platinum electrodes. [1]
- (ii) Draw a diagram of apparatus that would allow the gases produced in the reaction in (c) (i) to be collected separately. Annotate your diagram to show the polarity of each electrode and the names and relative volumes of each gas. [3]
- (d)



Two copper strips **X** and **Y** are placed in an aqueous solution of copper(II) sulfate and electrolyzed for a certain time. **X** was then dried and weighed.

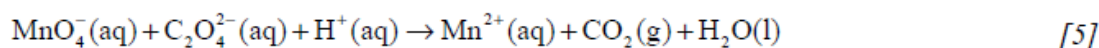
- (i) State and explain what would happen to the mass of **X**. [3]
- (ii) State **two** ways in which the change in the mass of **X** could be increased. [2]

HL B 02s

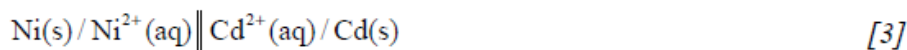
DATA BOOKLET

6. (a) The electrolysis of sodium chloride can be carried out with the sodium chloride in the molten state or as a concentrated solution in water.
- (i) Sketch a diagram showing how molten sodium chloride could be electrolysed in a beaker. Show clearly the polarity of the electrodes, the direction of the electron flow in the connecting wires and the products at **each** electrode. Write equations to show the formation of the product at **each** electrode. [5]
- (ii) When aqueous sodium chloride solution is electrolysed, the product at one of the electrodes is different. Name this product and explain why it is formed instead of the product in (i). [3]
- (iii) In the electrolysis of molten sodium chloride, a current of 5 A is passed for 1 hour. Calculate the mass of the product formed at the negative electrode. [3]
- (b) Ethanedioate ions, $\text{C}_2\text{O}_4^{2-}$, can be oxidised to carbon dioxide by acidified dichromate(VI) ions.
- (i) Deduce the half-equation for the oxidation of ethanedioate ions. Deduce the oxidation number of carbon in ethanedioate ions and in carbon dioxide, and use your values to explain why carbon is oxidised in this reaction. [4]
- (ii) Using information from the Data Booklet, write the equation for the reaction between ethanedioate ions and dichromate(VI) ions in acid solution. [2]
- (c) An electrochemical cell is constructed from two half-cells connected by a high-resistance voltmeter. One half-cell contains nickel in a solution of nickel nitrate, and the other contains silver in silver nitrate solution.
- (i) State the conditions which must apply to the solutions for the measurements made to be described as *standard*. [2]
- (ii) Outline how the two half-cells must be connected before any voltage readings can be made. [2]
- (iii) Assuming that standard conditions apply, calculate the cell potential using information from the Data Booklet. Write the shorthand notation for the cell, including state symbols, and give the equation for the reaction occurring in the cell. [4]

7. (a) Redox equations may be balanced using changes in oxidation number. For the following redox equation calculate the oxidation number of manganese and carbon. Use these values to balance the equation.

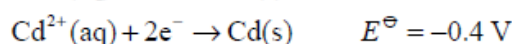
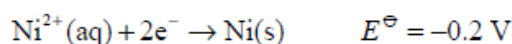


- (b) (i) Draw a cell diagram for the cell formed by connecting the following **standard** half-cells:

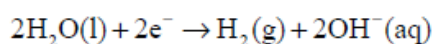


- (ii) Describe the key features of the standard hydrogen electrode. [3]

- (c) Given:



- (i) Write an equation for the reaction in **each** half-cell, identify the species which is oxidised **and** the oxidising agent. [4]
- (ii) On the diagram of this cell drawn in (b) (i), label the anode (**A**), and show, with an arrow, the direction of electron flow in the external circuit. [2]
- (iii) For the overall cell, calculate its voltage and state the sign of ΔG . [2]
- (d) An aqueous solution of silver nitrate is electrolysed. Predict the product formed at each electrode. [2]
- (e) A membrane cell is used to electrolyse aqueous sodium chloride. Hydrogen and sodium hydroxide are produced according to the following equation:



A current of 20 A is passed through the solution for 5 hours. Calculate the number of moles of OH^- produced, and the mass of sodium hydroxide formed. [4]

HL B 00w

5

- (e) Describe and explain the redox reactions of Cl_2 , Br_2 and I_2 with Cl^- , Br^- and I^- ions. [5]

HL B .99w

TABLE 15

7. (a) The properties of an unknown metal **M** and its ions are being investigated.
- (i) Assuming that there is a supply of the metal and its soluble sulfate, describe how the standard electrode potential, $M^{2+}(aq) / M(s)$, could be determined.
(You may answer by drawing a fully-labelled diagram if you wish.) [8]
- (ii) **M** also forms an oxyanion, MO_3^- , that is readily converted to M^{2+} in acid solution. Calculate the oxidation number of **M** in the MO_3^- ion. State and explain whether $MO_3^- \rightarrow M^{2+}$ is an oxidation or a reduction process. Write a balanced half-equation for this reaction. [3]
- (iii) The standard electrode potential for the reaction in (a) (ii) is found to be +1.2 V. Using Table 15 in the Data Booklet explain how this value could be used to predict whether MO_3^- would oxidise Br^- ions to Br_2 or whether Br_2 would oxidise colourless M^{2+} to MO_3^- . Write a balanced equation for the reaction that would be expected to occur. How could this prediction be tested experimentally? (Give the observations you would expect for both a positive and negative result.) Suggest a reason why this prediction might not be realised in practice. [8]
- (b) Electrolysis of molten potassium bromide and aqueous potassium bromide produces some products which are the same, and others which are different. Write balanced equations for the electrolysis in each case and account for the different products formed. [6]

IB HL 9 EQ Paper 2 s99 to s13 incl W Mark Scheme

HL SECTION A 10w

1

- (e) (i) the more negative the electrode potential the more negative the enthalpy change / the more exothermic the reaction / *OWTTE*; [1]
Accept electrode potential proportional to the enthalpy change.
Do not accept greater enthalpy change.
- (ii) any line (straight or curve) with positive gradient;
Accept if curve becomes flat for electrode potentials above 0.25.
- which passes through $\Delta H = 0$ at standard electrode potentials between 0.25 and 0.50; [2]
Accept either positive or zero enthalpy change for electrode potentials greater than value of copper (in range 0.25 to 0.50).

HL SECTION A 10s

1. (a) (i) *Copper:*
0 to +2 / increases by 2 / +2 / 2+;
Allow zero/nought for 0.
- Nitrogen:*
+5 to +4 / decreases by 1 / -1 / 1-;
Penalize missing + sign or incorrect notation such as 2+, 2⁺ or II, once only. [2]
- (ii) *nitric acid/HNO₃ / NO₃⁻/nitrate;* [1]
Allow nitrogen from nitric acid/nitrate but not just nitrogen.
- (b) (i) 0.100 × 0.0285 ;
2.85 × 10⁻³ (mol); [2]
Award [2] for correct final answer.
- (ii) 2.85 × 10⁻³ (mol); [1]
- (iii) (63.55 × 2.85 × 10⁻³) = 0.181 g ; [1]
Allow 63.5.
- (iv) $\left(\frac{0.181}{0.456} \times 100 =\right)$ 39.7 %; [1]
- (v) $\left(\frac{44.2 - 39.7}{44.2} \times 100 =\right)$ 10/10.2 %; [1]
Allow 11.3 % i.e. percentage obtained in (iv) is used to divide instead of 44.2 %.
- (c) (i) 1s²2s²2p⁶3s²3p⁶3d⁹; [1]
Do not allow [Ar]3d⁹.
- (ii) d orbitals are split;
(3d) electrons move between orbitals **and** absorb light/energy / complementary colour is transmitted when energy absorbed by d electrons moving; [2]
Accept levels instead of orbitals.

HL SECTION A 09w

1. (a) $\text{MnO}_4^- (\text{aq}) + 5\text{Fe}^{2+} (\text{aq}) + 8\text{H}^+ (\text{aq}) \rightarrow \text{Mn}^{2+} (\text{aq}) + 5\text{Fe}^{3+} (\text{aq}) + 4\text{H}_2\text{O} (\text{l})$ [2]
Award [2] if correctly balanced.
Award [1] for correctly placing H⁺ and H₂O.
Award [1 max] for correct balanced equation but with electrons shown.
Ignore state symbols.
- (b) Fe²⁺ / iron(II); [1]
Do not accept iron.
- (c) $n = 2.152 \times 10^{-2} \times 2.250 \times 10^{-2}$;
4.842 × 10⁻⁴ (mol); [2]
Award [1] for correct volume
Award [1] for correct calculation.

HL SECTION A 09w

4. (a) electrolytic cell converts electrical energy to chemical energy **and** voltaic cell converts chemical energy to electrical energy / electrolytic cell uses electricity to carry out a (redox) chemical reaction **and** voltaic cell uses a (redox) chemical reaction to produce electricity / electrolytic cell requires a power supply **and** voltaic cell does not;

electrolytic cell involves a non-spontaneous (redox) reaction **and** voltaic cell involves a spontaneous (redox) reaction;

in an electrolytic cell, cathode is negative and anode is positive **and** *vice-versa* for a voltaic cell / electrolytic cell, anode is positive and voltaic cell, anode is negative / electrolytic cell, cathode is negative and voltaic cell, cathode is positive;

voltaic cell has two separate solutions **and** electrolytic cell has one solution / voltaic cell has salt bridge and electrolytic cell has no salt bridge;

electrolytic cell, oxidation occurs at the positive electrode/anode **and** voltaic cell, oxidation occurs at the negative electrode/anode/*vice-versa*; [2 max]

If descriptions are reversed for electrolytic and voltaic cell, penalize first marking point but award second marking point as ECF.

- (b) (i) $2\text{Al}(s) + 3\text{Ni}^{2+}(\text{aq}) \rightarrow 2\text{Al}^{3+}(\text{aq}) + 3\text{Ni}(s)$; [2]

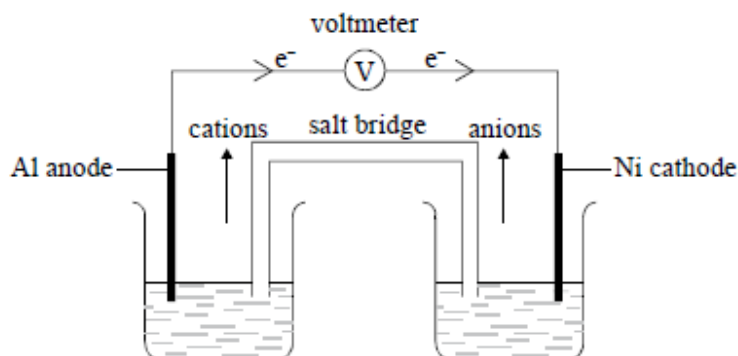
Correct reactants and products, award [1]

Balancing award [1].

Ignore state symbols and equilibrium sign.

- (ii) (+)1.40 (V); [1]

- (iii) aluminium anode/negative electrode;
nickel cathode/positive electrode;
electron movement from Al to Ni;
correct movement of cations and anions through salt bridge; [4]
If electron movement shown correctly but not labelled, award the mark



4. (a) (i) $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$; [1]
- (ii) $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$; [1]
- (iii) $\text{MnO}_4^- + 5\text{Fe}^{2+} + 8\text{H}^+ \rightarrow \text{Mn}^{2+} + 5\text{Fe}^{3+} + 4\text{H}_2\text{O}$; [1]

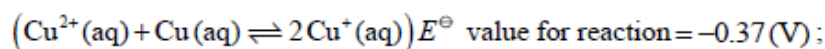
Accept e instead of e⁻.

- (b) (i) amount of $\text{MnO}_4^- = \frac{11.6}{1000} \times 0.0200 = 2.32 \times 10^{-4} \text{ mol}$; [1]
- (ii) amount of $\text{Fe}^{2+} = 5 \times 2.32 \times 10^{-4} = 1.16 \times 10^{-3} \text{ mol}$; [1]
- (iii) mass of $\text{Fe}^{2+} = 55.85 \times 1.16 \times 10^{-3} = 6.48 \times 10^{-2} \text{ g}$;
percentage of Fe^{2+} in tablet = $\frac{6.48 \times 10^{-2}}{1.43} \times 100 = 4.53\%$; [2]

HL SECTION A 07w

4. (a) (i) $\text{I}^- = -1/1-$ and $\text{IO}_3^- = +5/5+$; [1]
*Both answers needed for [1] mark,
Signs must be included
Do not accept Roman numerals*
- (ii) *oxidation*
 I^- (to I_2), increase in oxidation number / loss of electron(s);
reduction
 IO_3^- (to I_2), decrease in oxidation number / gain of electron(s); [2]
- (iii) $5\text{I}^-(\text{aq}) + \text{IO}_3^-(\text{aq}) + 6\text{H}^+(\text{aq}) \rightarrow 3\text{I}_2(\text{aq}) + 3\text{H}_2\text{O}(\text{l})$; [2]
*Award [2] if correctly balanced
Award [1] for correct reactants and products.
States not required for mark.
[1 max] if HCl on left and Cl⁻ on right side.*

- (b) $\text{Cu}^{2+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Cu}^+(\text{aq})$; $E^\ominus = (+)0.15(\text{V})$, $\text{Cu}(\text{s}) \rightleftharpoons \text{Cu}^+(\text{aq}) + \text{e}^-$; $E^\ominus = -0.52(\text{V})$ /
(+)0.15 and -0.52(V);
No mark if 0.34 or 0.52 quoted, but then ECF



Award [2] for correct E_{rxn}^\ominus even if equations are not given, states not required.

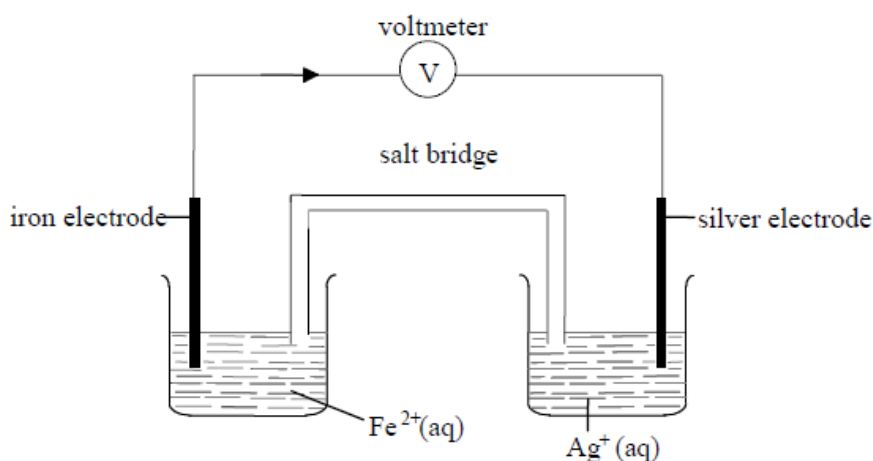
(negative value means) not spontaneous;

Allow ECF: if positive value, can score third mark for 'spontaneous'.

[3]

HL SECTION A 07s

2. (a) $\text{Fe} + 2\text{Ag}^+ \rightarrow \text{Fe}^{2+} + 2\text{Ag}$; [1]
Ignore state symbols.
Accept $\text{Fe} + 3\text{Ag}^+ \rightarrow \text{Fe}^{3+} + 3\text{Ag}$
- (b) the potential difference / EMF / Voltage between a standard half-cell and standard hydrogen electrode / *OWTTE*; [1]
- (c) (+) 1.24 (V); [1]
ECF from (a).
- (d) electron flow indicated on wires; [1]
ECF from (a).



HL SECTION A 07s

3. (a) *oxidizing agent: (acidified) potassium permanganate(VII) / (H⁺) and MnO₄⁻ and reducing agent: Sn²⁺;* [1]
Both oxidizing agent and reducing agent required for [1].
- (b) $5\text{Sn}^{2+} + 2\text{MnO}_4^- + 16\text{H}^+ \rightarrow 5\text{Sn}^{4+} + 2\text{Mn}^{2+} + 8\text{H}_2\text{O}$; [1]

HL SECTION A 06s

4. (a) (i) (a species that) gains electrons (from another species) / causes electron loss; [1]
- (ii) changes by 3;
 reduced because its oxidation number decreased / +6 → +3 / 6+ → 3+ / it has gained electrons; [2]
- (b) (i) $\text{C}_6\text{H}_8\text{O}_6 \rightarrow \text{C}_6\text{H}_6\text{O}_6 + 2\text{H}^+ + 2\text{e}^-$; [1]
- (ii) $\text{C}_6\text{H}_8\text{O}_6 + 2\text{Fe}^{3+} \rightarrow \text{C}_6\text{H}_6\text{O}_6 + 2\text{H}^+ + 2\text{Fe}^{2+}$; [1]

HL SECTION A 03w

5. (a) (i) oxidation half-reaction: $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$;
 reduction half-reaction: $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$ / $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$;
Award [1] only if equations are interchanged. [2]
States not required.

- (ii) Na has high $E_{\text{red}}^\ominus / \text{Na}^+$ not readily reduced (in comparison to H_2O) / if
 formed, Na would (immediately) react with water to form Na^+ [1]

- (b) $\text{H}_2(\text{g})$ and $\text{O}_2(\text{g})$ / *accept names*;
 $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$ / $4\text{OH}^- \rightarrow 2\text{H}_2\text{O} + \text{O}_2 + 4\text{e}^-$;
 water is oxidized (instead of the halide);
 since E_{ox}^\ominus for F^- is very negative / E_{red}^\ominus for F_2 is very high;
Accept answer based on oxidizing / reducing strengths. [4]

HL SECTION A 02w

4. (a) allows movement of **ions** between the solutions (thus completing the circuit) [1]

- (b) (i) $\text{Fe}^{2+} + 2\text{e}^- \rightleftharpoons \text{Fe} \quad -0.44 \text{ V}$
 $\text{Fe}^{3+} + \text{e}^- \rightleftharpoons \text{Fe}^{2+} \quad 0.77 \text{ V}$ [1]
Both needed for mark.

E^\ominus of cell = $0.77 + 0.44 = 1.21 \text{ V}$ [1]
Correct answer without working, award [2]. -1.21 V, award [1].

- (ii) A: $\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$;
 B: $\text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^-$; [2]
A and B equations interchanged, award [1].

- (iii) B to A / Fe to Pt [1]
Allow for ECF.

HL SECTION A 01w

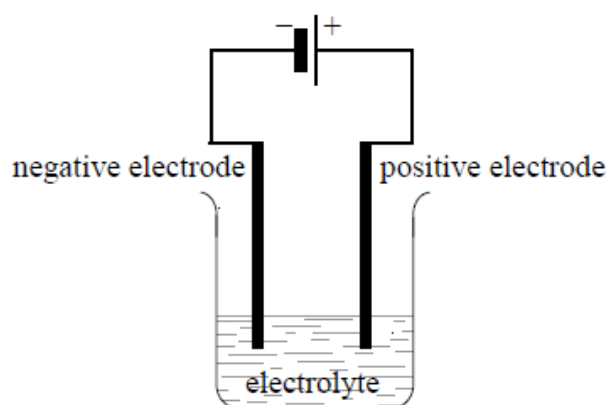
4. (a) Anode: $\text{Cu(s)} \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^-$ (state symbols needed) [1]
 Cathode: $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu(s)}$ (state symbols needed) [1]
- (b) Remains the same as $[\text{Cu}^{2+}]$ does not change (reasoning needed) [1]
 (ECF if reaction in 4 (a) is incorrect)
- (c) $2\text{H}_2\text{O(l)} \rightarrow \text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^-$ (states not required) [1]
 (Award [1] for balanced half equation and [1] for O_2 and H^+ as products.) [2]
- (d) The colour intensity decreases / changes [1]
 The pH decreases / changes [1]
- (e) $C = At = 0.180 \text{ A} \times 1210 \text{ s} = 218 \text{ C}$ [1]
 Number of Faradays, $F = \frac{218 \text{ C}}{96480 \text{ C}} (\text{mole}^-) = 2.26 \times 10^{-3} \text{ mole}^-$ [1]
 2 mole^- produce 1 mol Cu = 63.35 g Cu
 thus $2.26 \times 10^{-3} \text{ mole}^-$ gives $0.50 \times 63.35 \times 2.26 \times 10^{-3} \text{ g} = 0.0715 \text{ g Cu}$ [1]

HL SECTION A 00s

4. (a) Reducing agent donates/loses electrons / OWTTE. [1]
- (b) Current flow: $\text{Al} \rightarrow \text{Ni}$ [1]
 $\text{Al} \ominus, \text{Ni} \oplus$ (both) [1]
- (c) $2\text{Al} + 3\text{Ni}^{2+} \rightarrow 3\text{Ni} + 2\text{Al}^{3+}$ [2]
 (Award [1] for correct species on correct sides of equation and [1] for correct coefficients, even if equation reversed.)
 $E^\ominus = +1.43 \text{ V}$ [2]
 (Award [1] each for sign and value. Allow -1.43 V if equation reversed – ECF principle. If signs not appropriate but value correct, award [1]. If E^\ominus values added, award [0].)
- (d) Seconds = 2×3600 OR 7200 [1]
 Coulombs = $8 \times 2 \times 3600$ OR 57600 [1]
 $\div F = \frac{57600}{96480}$ OR 0.5970 [1]
 $\div 6 = \frac{57600}{96480 \times 6}$ [1]
 Answer = $0.09950 = 0.10(00)$ [1]
 (Correct answer with no working, [4 max].)

HL B 13s

(c) (i)



clear diagram containing all elements (power supply, connecting wires, electrodes, container and electrolyte);

Accept power supply if shown as conventional long/short lines (as in diagram above) or clearly labelled DC power supply.

labelled positive electrode/anode **and** negative electrode/cathode;

Accept positive and negative by correct symbols near power supply.

labelled electrolyte/ $\text{FeBr}_2(\text{l})/\text{FeBr}_2(\text{aq})$;

State must be included for FeBr_2 .

[3]

(ii) *Electrolyte*: positive ions/cations move to negative electrode/cathode **and** negative ions/anions to positive electrode/anode;

Conductors: electrons flow from negative pole of battery to positive pole of battery / *OWTTE*;

Look at diagram in (i) for possible clarification of electron flow.

[2]

*Award [1 max] for "electrons in wire/external circuit **and** ions in solution".*

(iii) *Negative electrode/cathode*:

H_2 ;

$E^\ominus(\text{H}_2)$ is less negative than $E^\ominus(\text{Fe})$ / Fe is more reactive than H_2 / H_2 is lower in reactivity series / H^+ more easily reduced than Fe^{2+} / *OWTTE*;

Positive electrode/anode:

O_2 ;

$E^\ominus(\text{O}_2)$ is less positive than $E^\ominus(\text{Br}_2)$ / in a dilute Br^- solution $\text{OH}^-/\text{H}_2\text{O}$ is preferably discharged / *OWTTE*;

Award [3 max] if electrodes reversed or omitted.

[4]

(iv) Br_2 ;

Accept Fe .

[1]

(v) $2\text{Br}^- \rightleftharpoons \text{Br}_2 + 2\text{e}^-$ shifts to the right;

Accept similar reason for Fe .

[1]

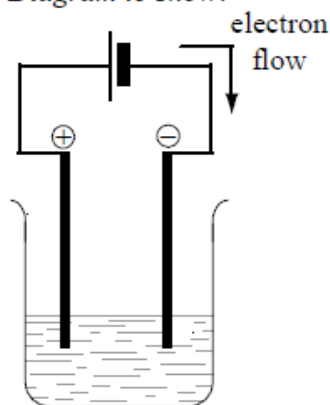
4. (a) (i) *Oxidation*: loss of electrons **and** *Reduction*: gain of electrons; [1]

(ii) As_2O_3 : +3;
 NO_3^- : +5;
 H_3AsO_3 : +3;
 N_2O_3 : +3; [4]
Penalize incorrect notation e.g. III, V, 3+, 5+, 3, 5 once only.

(iii) *Oxidizing agent*: substance reduced / removes electrons from another substance / causes some other substance to be oxidized / **OWTTE and Reducing agent**: substance oxidized / gives electrons to another substance / causes some other substance to be reduced / **OWTTE**; [1]
Accept Oxidizing agent: electron/ e^- acceptor / causes oxidation / oxidation number/state decreases and Reducing agent: electron/ e^- donor / causes reduction / oxidation number/state increases.

(iv) $As_2O_3(s) + 2NO_3^-(aq) + 2H^+(aq) + 2H_2O(l) \rightarrow 2H_3AsO_4(aq) + N_2O_3(aq)$
correct coefficients for As_2O_3 , H_3AsO_4 **and** NO_3^- , N_2O_3 ;
correct balanced equation;
Ignore state symbols.
M1 must be correct to award M2.
*Oxidizing agent: $NO_3^-(aq)$ /nitrate **and** Reducing agent: $As_2O_3(s)$ /arsenic(III) oxide;*
Accept $HNO_3(aq)$ /nitric acid.
Accept arsenic oxide.
Species must be fully correct to score M3.
Ignore state symbols. [3]

(b) (i) *Diagram to show:*



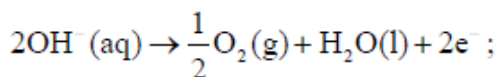
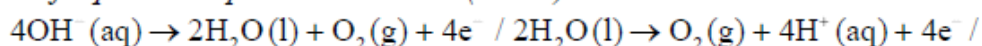
Labels are not required.
one container, two electrodes, battery (and electrolyte);

Allow + and - for representation of battery (could be long and short lines for example) but M1 is not scored if a voltmeter/V if shown or labelled.
Ignore designation of electrodes (e.g. do not penalize Cu electrodes).

correct direction of electron flow (from negative pole to positive pole); [2]
Allow arrow without stating e^- explicitly.
To score M2, the polarity of the battery or the cathode and anode must be shown.

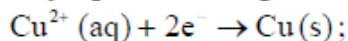
If a voltmeter/V is shown, M1 is not awarded but M2 may be scored if the cathode and anode are identified with the correct direction of electron flow.

(ii) *Half-equation at positive electrode (anode):*



Allow e instead of e⁻.

Half-equation at negative electrode (cathode):



Allow e instead of e⁻.

Award [1 max] for M1 and M2 if correct equations are given but at wrong electrodes.

Penalize \rightleftharpoons once only in (b)(ii) and (b)(iii).

correct state symbols in **all** equations;

M3 can only be scored if the correct species are given in M1 and M2 (i.e. do not award ECF from M1 and M2 for incorrect species).

Observation at positive electrode (anode):

bubbles / gas;

Award mark for observation even if type of gas is incorrect (e.g. hydrogen).

Observation at negative electrode (cathode):

red/brown/copper/metal (deposit);

Allow mass increases / gets thicker/larger / OWTTE.

Change in colour (in any) of the solution:

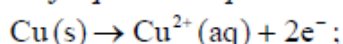
solution loses blue colour/becomes paler;

Change in acidity (if any) of the solution:

becomes (more) acidic / pH decreases;

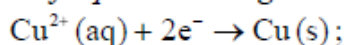
[7]

(iii) *Half-equation at positive electrode (anode):*



Allow e instead of e⁻.

Half-equation at negative electrode (cathode):



Allow e instead of e⁻.

Award [1 max] for M1 and M2 if correct equations are given but at wrong electrodes.

Penalize \rightleftharpoons once only in (b)(ii) and (b)(iii).

correct state symbols in **all** equations;

M3 can only be scored if the correct species are given in M1 and M2 (i.e. do not award ECF from M1 and M2 for incorrect species).

Observation at positive electrode (anode):

(slowly) dissolves / OWTTE;

Allow mass decreases / gets smaller/thinner / OWTTE;

Accept impurities deposited under positive electrode/anode.

Observation at negative electrode (cathode):
red/brown/copper/metal deposit;
Allow mass increases / gets thicker/larger / OWTTE.

Change in colour (if any) of the solution:
blue colour remains;
Allow no change or solution remains same colour.

Change in acidity (if any) of the solution:
solution does not become (more) acidic / no change / OWTTE; [7]

HL B 12s

6

(e) (i) Br_2 : 0
 HBr : -1
 $HOBr$: +1 [2]
Award [2] for three correct.
Award [1] for any two correct.

(ii) bromine is oxidized **and** reduced / disproportionation; [1]

(f) (i) F_2 /fluorine; [1]
Do not allow F.

(ii) $50 \text{ (cm}^3\text{)} / 0.050 \text{ dm}^3$; [1]

HL B 11w

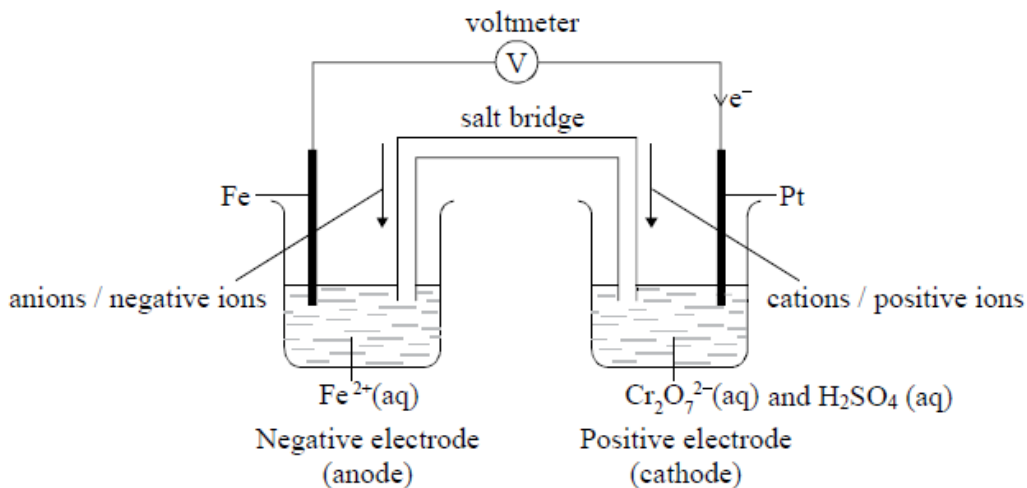
7. (a) *Oxidation:* increase in oxidation number **and** *Reduction:* decrease in oxidation number / OWTTE; [1]

(b) Cr_2O_3 :
chromium(III) oxide;
Do not award mark for chromium oxide.
 CrO_3 :
chromium(VI) oxide; [2]
Do not award mark for chromium oxide.
Do not award any marks if chromium oxide without Roman numerals is given for both.

(c) (i) substance reduced / causes other substance to be oxidized / increase oxidation number of another species / gains electrons / OWTTE; [1]

(ii) *Oxidizing agent:*
 $Cr_2O_7^{2-}$ / dichromate (ion);
 $Cr_2O_7^{2-} (aq) + I^- (aq) + 8H^+ (aq) \rightarrow 2Cr^{3+} (aq) + IO_3^- (aq) + 4H_2O(l)$ [3]
Award [1] for coefficients: $Cr_2O_7^{2-} (aq)$, $I^- (aq)$, $2Cr^{3+} (aq)$, $IO_3^- (aq)$.
Award [1] for coefficients: $8H^+ (aq)$, $4H_2O(l)$.
Award [1 max] if coefficients of reactants only correct i.e. $Cr_2O_7^{2-}$, I^- and $8H^+$.
Award [1 max] if coefficients of products only correct i.e. $2Cr^{3+}$, IO_3^- and $4H_2O$.
Award [1 max] for correct reactants and products.
Ignore state symbols.

(d) (i)



Voltaic cell showing:

labelled positive electrode (cathode) **and** negative electrode (anode);

direction of electrons in external circuit **and** direction of ions in salt bridge;

Award mark if correct direction of electrons is indicated but e^- not labelled in external circuit.

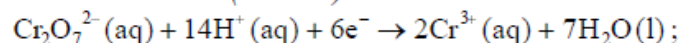
Allow e instead of e^- .

Cations/positive ions and anions/negative ions must be identified in salt bridge.

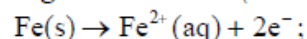
Allow correct movement of ions in electrolyte instead of movement of ions in salt bridge (e.g. Fe^{2+} from Fe at negative electrode/anode etc.).

If both movement of ions in salt bridge and movement of ions in electrolyte is given but one is incorrect do not award mark.

Positive electrode (cathode):



Negative electrode (anode):

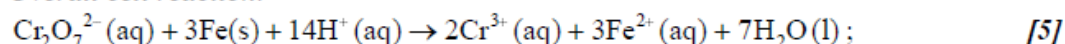


Penalize \rightleftharpoons once only.

Penalize once only if electrodes or equations reversed.

For both electrodes allow e instead of e^- .

Overall cell reaction:



Ignore state symbols throughout (d) (i).

(ii) potential under standard conditions relative to standard hydrogen electrode/SHE; [1]

Reference must be made to standard conditions.

Instead of standard conditions allow either solute concentration of $1 \text{ mol dm}^{-3}/1 \text{ M}/1 \text{ mol L}^{-1}$ or $100 \text{ kPa}/10^5 \text{ Pa}$ for gases.

Allow 1 bar for $100 \text{ kPa}/10^5 \text{ Pa}$.

Allow 1 atm/ $1.01 \times 10^5 \text{ Pa}$.

Allow voltage instead of potential.

(iii) (+)1.78 (V); [1]

(f) *Positive electrode (anode): chromium;
Allow lead/titanium/platinum/graphite.*

*Negative electrode (cathode): object to be plated;
Allow specific example here e.g. spoon.*

Electrolyte: $\text{Cr}^{3+}(\text{aq})$;

[3]

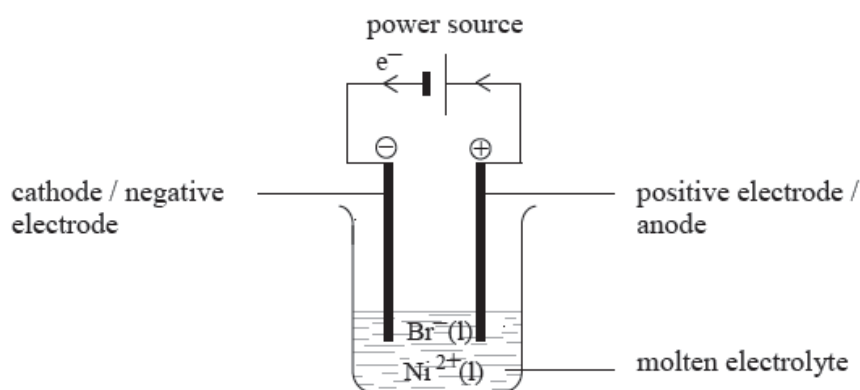
Allow (mixture of) $\text{Cr}^{3+}(\text{aq})$ and $\text{CrO}_4^{2-}(\text{aq})$ /chromate/chromic acid/ H_2CrO_4 .

Ignore state symbols.

Allow any soluble salt of Cr^{3+} .

HL B 11s

9. (a) (i)

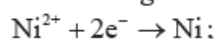


labelled polarities of positive **and** negative electrodes;
Electrodes can be labelled positive or negative or + and - signs can be used.

direction of electron flow;
e⁻ does not have to be labelled but arrow essential.

power source **and** molten electrolyte/Ni²⁺ (l) **and** Br⁻ (l)/NiBr₂ (l);
State symbol necessary for M3 unless molten electrolyte stated.
Power source does not need to be labelled if correct symbol used (i.e. short line and long line).

Cathode/negative electrode equation:



Anode/positive electrode equation:



[5]

Accept balanced half-equation with one e⁻.

Award [1 max] for M4 and M5 if electrodes are not identified or if equations are given wrong way round or incorrectly labelled.

Penalize ⇌ once only in Q.9.

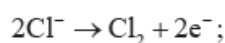
Allow e instead of e⁻.

Ignore state symbols for M4 and M5.

(ii) *Dilute sodium chloride:*



Concentrated sodium chloride:



[2]

Accept alternative balanced half-equations with correct number of electrons.

Award [1 max] if equations are given the wrong way round.

Award [2] if correct equations are written in order with dilute sodium chloride first and concentrated sodium chloride second but processes not stated explicitly.

Penalize ⇌ once only in Q.9.

Allow e instead of e⁻.

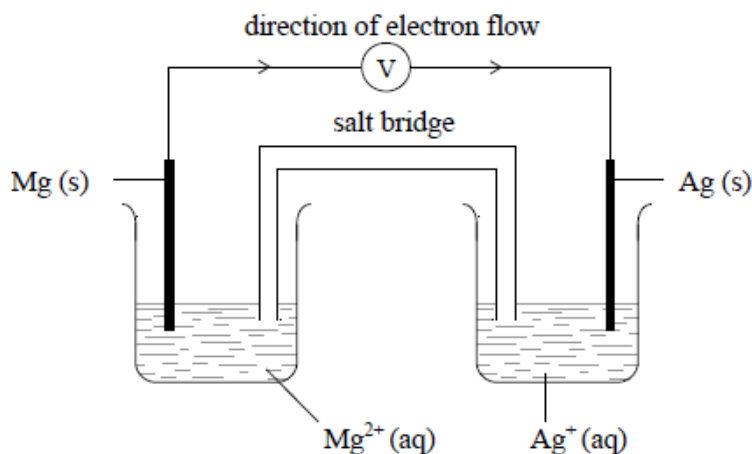
Ignore state symbols.

- (b) (i) $\text{Sn} + \text{Cu}^{2+} \rightarrow \text{Sn}^{2+} + \text{Cu}$; [1]
Ignore state symbols.
Penalize \rightleftharpoons once only in Q.9.
- (ii) $(0.34 - -0.14) = (+) 0.48 \text{ V}$; [1]
- (iii) 1.0 mol dm^{-3} solutions **and** $25^\circ\text{C} / 298 \text{ K}$; [1]
- (c) Cd^{2+} is a stronger oxidizing agent than H_2O **and** will be displaced to produce Cd / *OWTTE*;
 Cr^{2+} is a weaker oxidizing agent than H_2O **and** H_2 will displace in preference to Cr / *OWTTE*; [2]
*Award [1 max] for stating Cd^{2+} stronger oxidizing agent than H_2O **and** Cr^{2+} weaker oxidizing agent than H_2O / *OWTTE*.*
- (d) (i) Ni;
 only requires 2 mol of e^- for each mol of Ni / Sn requires 4 mol of e^- / Cr requires 3 mol of e^- / Ni^{2+} needs least number of e^- to produce 1 mol of Ni metal;
Allow e instead of e^- .
- cathode / negative electrode; [3]
Do not award M3 for "metal deposited at cathode where oxidation occurs".
- (ii) temperature of solution;
 $[\text{Sn}^{4+}]$;
 surface area/size of electrode;
 material of electrodes; [2 max]
Do not allow nature of electrodes.
- (e) (i) NH_3 : -3;
 NO : +2;
 N_2 : 0; [3]
Penalize incorrect notation such as 3-, III, 2+, 2, II once only.
- (ii) *Oxidation:*
 $2\text{NH}_3 \rightarrow \text{N}_2 + 6\text{H}^+ + 6e^-$;
Reduction:
 $2\text{NO} + 4\text{H}^+ + 4e^- \rightarrow \text{N}_2 + 2\text{H}_2\text{O}$;
Award [1 max] for M1 and M2 if redox processes are not identified or if equations are given wrong way round.
Penalize \rightleftharpoons once only in Q.9.
Allow e instead of e^- .
Ignore state symbols.
- Oxidizing agent: NO;* [3]
Allow either formula or name.
- (iii) NH_3 /ammonia (in excess by) **and** $10 \text{ (dm}^3\text{)}$;
 $25.0 \text{ (dm}^3\text{)}$; [2]

- (b) (i) contains ions which are free to move (only) in molten state;
 Mg^{2+} move to cathode/negative electrode and Cl^- move to anode/positive electrode / *OWTTE*; [2]
- (ii) anode/positive electrode;
 $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$ / $\text{Cl}^- \rightarrow \frac{1}{2}\text{Cl}_2 + \text{e}^-$; [2]
Accept e instead of e^- .
Do not accept $\text{Cl}^- \rightarrow \text{Cl} + \text{e}^-$.
Ignore state symbols.
- (iii) magnesium has large negative electrode potential / E^\ominus ;
 reduction of $\text{H}_2\text{O}/\text{H}^+$ to H_2 has less negative electrode potential;
 Mg^{2+} not readily reduced (in comparison to H_2O);
 if formed, magnesium would (immediately) react with water to form Mg^{2+} ;
 magnesium more reactive than hydrogen; [1 max]
Do not accept Mg too reactive.

HL B 10s

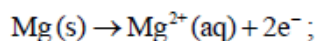
6. (a) (i)



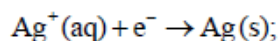
correctly labelled electrodes **and** solutions;
labelled salt bridge;
voltmeter;
Allow bulb or ammeter.

direction of electron flow; [4]

(ii) *Oxidation:*



Reduction:



[2]

Ignore state symbols.

Award [1 max] if equations not labelled reduction or oxidation or labelled the wrong way round.

Allow e instead of e⁻.

Penalize equilibrium sign or reversible arrows once only in parts (a) (ii) and (d) (ii).

(iii) $+0.80 - (-2.37) = 3.17 \text{ V}$

correct data;

answer with unit;

[2]

Award [1] for -3.17 V or correct working of wrong values.

(b) (i) $\text{Cd}/\text{Cd(s)}$;

[1]

Do not allow Cd^{2+} .

(ii) $5\text{Cd(s)} + 2\text{MnO}_4^{-}(\text{aq}) + 16\text{H}^{+}(\text{aq}) \rightarrow 5\text{Cd}^{2+}(\text{aq}) + 2\text{Mn}^{2+}(\text{aq}) + 8\text{H}_2\text{O(l)}$

correct reactants and products;

correct balancing of this equation;

[2]

Ignore state symbols.

(c) *Accept suitable diagram with the following indicated:*

Pt electrode;

$[H^+(aq)] = 1 \text{ mol dm}^{-3} / 0.5 \text{ mol dm}^{-3} \text{ H}_2\text{SO}_4$;

H_2 gas;

at 1 atm / $1.01 \times 10^5 \text{ Pa}$;

Do not award mark for pressure if no hydrogen gas given.

298 K / 25°C ;

[4 max]

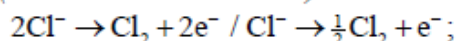
(d) (i) sodium chloride crystals consist of ions in a (rigid) lattice / ions cannot move (to electrodes) / *OWTTE*;
when melted ions free to move / ions move when potential difference/voltage applied;

[2]

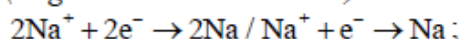
(ii) positive sodium ions/ Na^+ move to the negative electrode/cathode **and** negative chloride ions/ Cl^- move to the positive electrode/anode;

electrons are released to positive electrode/anode by negative ions and accepted from negative electrode/cathode by positive ions / reduction occurs at the negative electrode/cathode **and** oxidation occurs at the positive electrode/anode / Na^+ ions are reduced **and** Cl^- ions are oxidized;

(Positive electrode/anode):



(Negative electrode/cathode)



[4]

Award [1 max] if equations not labelled or labelled wrong way round.

Allow e instead of e^- .

Penalize equilibrium sign or reversible arrows once only in parts (a) (ii) and (d) (ii).

(iii) *Products:*

oxygen at positive electrode **and** hydrogen at negative electrode;

moles of Mg = 0.5 / mole ratio of $\text{O}_2:\text{H}_2$ is 1:2;

Can be implied by calculation.

$$\text{mass oxygen} = \left(\frac{1}{2} \times \frac{12.16}{24.31} \times 32.00 = \right) 8.00 \text{ g} ;$$

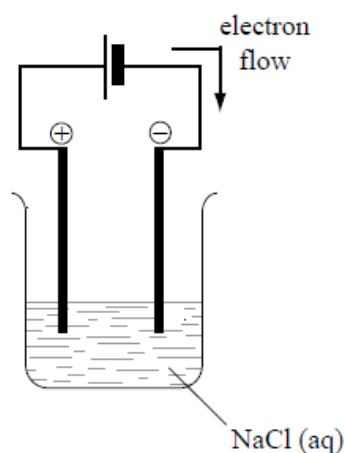
$$\text{mass hydrogen} = \left(\frac{12.16}{24.31} \times 2.02 = \right) 1.01 \text{ g} ;$$

[4]

Do not apply SD rule here.

7. (a) (i) the voltage obtained when the half-cell is connected to the standard hydrogen electrode;
under standard conditions of 298 K **and** 1 mol dm⁻³ solutions;
electrons flow (in the external circuit) from the half-cell to the hydrogen electrode / the metal in the half-cell is above hydrogen in the ECS / Fe is a better reducing agent than H₂ / Fe is oxidised more readily than H₂; [3]
- (ii) -0.28 V ; [1]
- (iii) Co²⁺/cobalt(II) ion; [1]
- (iv) $2\text{Al} + 3\text{Fe}^{2+} \rightarrow 3\text{Fe} + 2\text{Al}^{3+}$; [2]
Award [1] for correct reactants and products and [1] for correctly balanced, ignore states.
Do not accept \rightleftharpoons
- (v) to complete the electrical circuit / *OWTTE*;
by allowing the movement of ions; [2]
- (b) (i) +2; [1]
- (ii) +3; [1]
- (iii) +2; [1]
Only penalize once if roman numerals are used or if written as 2+ or 3+.

- (c) (i) diagram to show:



- battery/source of electricity connected to two electrodes in the solution with positive **and** negative electrodes correctly labelled;
electrons/current flowing from the cell to the negative electrode;
labelled solution of sodium chloride; [3]
If the connecting wires to electrodes are immersed in the solution [1 max].
- (ii) Na⁺, H⁺/H₃O⁺, Cl⁻, OH⁻ [2 max]
All four correct [2], any three correct [1].

- (iii) hydrogen at (-)/cathode **and** oxygen at (+)/anode;
 $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$ / $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$;
 $4\text{OH}^- \rightarrow \text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^-$ / $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$; [3]
 Accept e instead of e⁻
 If electrodes omitted or wrong way round [2 max].

- (iv) Ratio of H₂ : O₂ is 2 : 1; [1]

- (d) (i) (-)/(cathode) $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$ / $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$;
 (+)/(anode) $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$; [2]
 Accept e instead of e⁻.
 If electrodes omitted or wrong way round [1 max]

- (ii) (-)/(cathode) $\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$;
 (+)/(anode) $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$; [2]
 Accept e instead of e⁻.
 If electrodes omitted or wrong way round [1 max].

HL B 08w

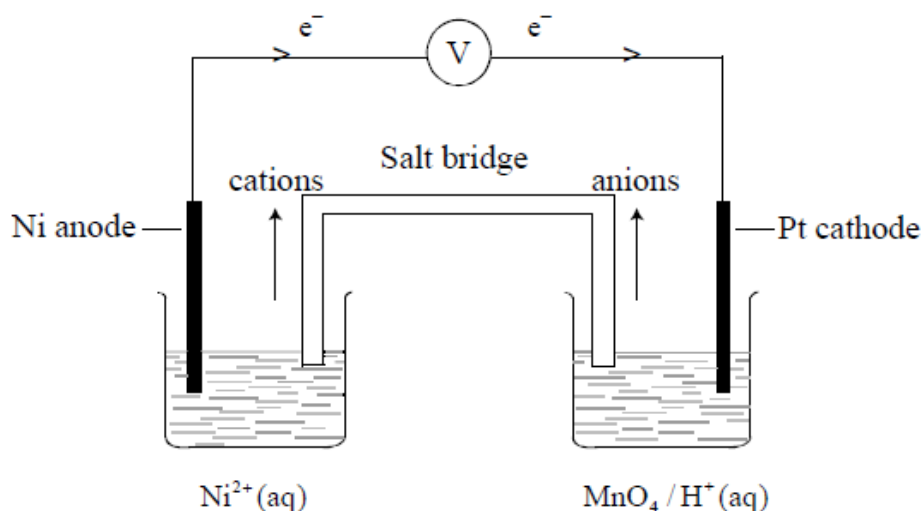
- 6
 (e) (i) $\text{Fe}^{2+} / \text{Ag}$;
 $2\text{Fe}^{2+} + \text{Br}_2 \rightarrow 2\text{Fe}^{3+} + 2\text{Br}^-$ / $2\text{Ag} + \text{Br}_2 \rightarrow 2\text{Ag}^+ + 2\text{Br}^-$ /
 $2\text{Ag} + \text{Br}_2 \rightarrow 2\text{AgBr}$; [2]
 Ignore state symbols.
- (ii) $(E^\ominus = +0.80 - (-1.66) =) (+)2.46 \text{ V}$; [1]
- (iii) $(E^\ominus = +0.48 - 0.76 =) - 0.28 \text{ V}$; [1]

HL B 08s

8. (a) (i) $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^- / 4\text{OH}^- \rightarrow \text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^-$; [2]
bubbles (of colourless gas) evolved;
Do not accept oxygen evolved as an observation.
- (ii) $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$; [2]
brown deposit;
Accept copper deposited
- (iii) electrolyte becomes acidic / pH decreases; [2]
blue colour becomes lighter / concentration of Cu^{2+} decreases;
Do not accept colour change.
- (iv) $\text{Cu} = 1 \text{ mol}$ **and** $\text{O}_2 = \frac{1}{2} \text{ mol}$; [1]
- (v) copper(II) nitrate / $\text{Cu}(\text{NO}_3)_2$; [1]
Accept any soluble copper(II) salt e.g. copper(II) ethanoate
- (vi) copper anode/positive electrode dissolves / decrease in mass; [3]
copper cathode/negative electrode increases in mass / brown deposit on
cathode;
no change in colour of the electrolyte / no change in the concentration of Cu^{2+} /
no change in pH;
- (b) current/I **and** time/t; [1]

- (c) (i) 25°C and 1.0 mol dm^{-3} ; [1]
Accept 298 K or 1.0 M but not 1 atm/101kPa
- (ii) $2\text{MnO}_4^{-}(\text{aq}) + 16\text{H}^{+}(\text{aq}) + 5\text{Ni}(\text{s}) \rightarrow 2\text{Mn}^{2+}(\text{aq}) + 5\text{Ni}^{2+}(\text{aq}) + 8\text{H}_2\text{O}(\text{l})$
 correct reactants and products;
 balancing; [2]
Ignore if state symbols not written.
Do not accept electrons in the final equation.
- (iii) reducing agent Ni(s)/Ni/Nickel;
Do not accept Ni²⁺.
 change in oxidation number 5 / +7 \rightarrow +2 / decreases by 5; [2]
- (iv) +1.74 V; [1]
- (v) *A clear-labelled diagram consisting of:*
 nickel anode;
 platinum cathode;
 electron movement from Ni to Pt;
 salt bridge;
 correct movement of cations / anions through salt bridge; [5]

e.g.



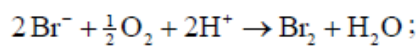
- (d) (voltaic cell)-spontaneous redox reaction / (electrolytic cell)-non-spontaneous redox reaction;
 (voltaic cell)-chemical energy converted to electrical energy / (electrolytic cell)-electrical energy converted to chemical energy; [2]
Accept anode is the positive electrode in an electrolytic cell and the negative electrode in a voltaic cell / cathode is the negative electrode in an electrolytic cell and the positive electrode in a voltaic cell.
For statements such as one electrolyte/two electrolyte and salt bridge/no salt bridge, award [1 max].

- (c) (i) (metal ions at) 1 mol dm^{-3} concentration;
 $25^\circ\text{C} / 298 \text{ K}$; [2]
Do not accept 1 atm pressure.
- (ii) $\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}$;
 $\text{Pb} \rightarrow \text{Pb}^{2+} + 2\text{e}^-$; [2]
No penalty for using e instead of e^- .
No penalty for \rightleftharpoons instead of \rightarrow .
- (iii) 0.15 V ;
 $10\text{Cl}^- + 2\text{MnO}_4^- + 16\text{H}^+ \rightarrow 5\text{Cl}_2 + 2\text{Mn}^{2+} + 8\text{H}_2\text{O}$; [3]
Ignore state symbols.
Correct reactants and products = [1]
Correct balancing = [1]
- (iv) not sufficiently good oxidizing agent / cell potential would have negative E^\ominus value; [1]

HL B 05w

7. (a) *at negative electrode (cathode)*
 $\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$;
- at positive electrode (anode)*
 $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$;
If both equations correct but electrodes incorrect or not stated, then deduct [1].
 electrons flow through the external circuit or wires;
 ions gain/lose electrons at electrodes/ions move to electrodes; [4]
- (b) 0.2 mol ;
 the $\text{Na} : \text{Cl}_2$ mol ratio is $2 : 1$ /correct reference to equations in (a); [2]
- (c) $0.1 \times \frac{1}{2} \times 5$;
 $= 0.25 \text{ mol}$; [2]
- (d) *negative electrode*
 hydrogen / H_2 ;
 $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$ / $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$;
- positive electrode*
 oxygen / O_2 ;
 $4\text{OH}^- \rightarrow 2\text{H}_2\text{O} + \text{O}_2 + 4\text{e}^-$ / $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$; [4]
- (e) (i) chlorine / Cl_2 gains electrons and is reduced;
bromide (ions) / Br^- loses electrons and is oxidized; [2]
Award [1] max if no mention of reduced and oxidized.
- (ii) *S in SO_2*
 $+4$;
S in H_2SO_4
 $+6$;
Award only [1] for 4+ and 6+ or 4 and 6.
 SO_2 oxidized because oxidation number (of sulfur) increases; [3]

- (f) the potential/voltage difference between the element and its ions;
(and) a hydrogen electrode;
under standard conditions/ion concentration at $1 \text{ mol dm}^{-3}/298 \text{ K}/25 \text{ }^\circ\text{C}$;
(+)1.23 (V);
 E^\ominus value more positive or less negative than bromine/bromide system /
 E^\ominus value of combined half-cells is positive/*OWTTE*;



Award [1] for all formulas correct, [1] for correct balancing.

Award [1] for correct equation reversed.

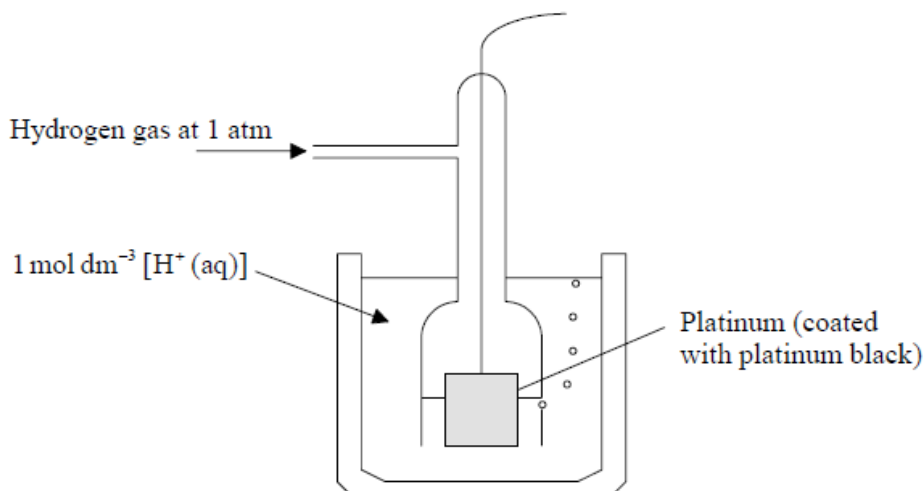
$$E^\ominus (= 1.23 - 1.09) = (+)0.14 \text{ (V)};$$

Ignore state symbols.

[8]

HL B 04w

9. (a)



Accept suitable diagram with the following indicated:

Pt electrode;

$1 \text{ mol dm}^{-3} [\text{H}^+(\text{aq})]$;

H_2 gas;

at 1 atm / $1.01 \times 10^5 \text{ Pa}$;

298 K / 25°C ;

[5]

(b) electron acceptor;

$\text{Fe}^{3+}(\text{aq})$ / iron(III) ions / Fe^{3+} ;

[2]

Do not accept iron / Fe^{2+} / iron ion.

(c) (i) (+)1.10;

[1]

(ii) $\text{Cu}^{2+}(\text{aq}) + \text{Zn}(\text{s}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu}(\text{s})$;

[2]

Award [1] for correct reactants and products from (c)(i),
and [1] for state symbols.

(d) (i) zinc;

zinc is more readily oxidised than iron and so protects it by reacting preferentially / *OWTTE* or tin is less readily oxidised than iron and so iron reacts preferentially / *OWTTE*;

[2]

(ii) charge on the ion discharged;

size / magnitude of the current;

time / duration of the electrolysis;

[3]

(iii) positive ions / cations in solution = $\text{H}^+(\text{aq})$, $\text{Zn}^{2+}(\text{aq})$;

$\text{H}^+(\text{aq})$ discharged preferentially;

[2]

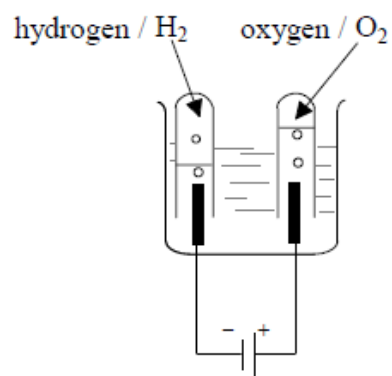
- (e) (i) salt bridge;
allows movement of ions between the solutions / to complete the circuit /
to maintain electrical neutrality; [2]
- (ii) A : $\text{Fe}^{3+}(\text{aq}) + \text{e}^{-} \rightarrow \text{Fe}^{2+}(\text{aq})$;
B : $\text{Cr}(\text{s}) \rightarrow \text{Cr}^{3+}(\text{aq}) + 3\text{e}^{-}$;
*Allow [2] for correct equation for the cell reaction if equations for A and B
are reversed.*
- $3 \text{Fe}^{3+}(\text{aq}) + \text{Cr}(\text{s}) \rightarrow 3 \text{Fe}^{2+}(\text{aq}) + \text{Cr}^{3+}(\text{aq})$; [4]
*Award [1] for correct reactants and products, with state symbols,
and [1] for correct balancing.*
- (iii) from B to A / from Cr to Pt / from right to left; [1]
Allow ECF from (ii).
- (iv) (+)1.51; [1]
Allow ECF from (ii).

HL B 04s

10. (a) (i) $[H^+] = 1 \text{ mol dm}^{-3}$;
 298 K / 25 °C ;
 1 atm / 101.3 or 101 kPa
Accept 100 kPa; [3]
- (ii) $E^\ominus (= 0.76 + 0.34) = (+)1.1(0) \text{ (V)}$;
 from zinc / Zn to copper / Cu;
 copper / Cu deposited / electrode becomes larger / thicker / heavier;
 zinc / Zn electrode becomes smaller / thinner / lighter;
 Cu^{2+} solution becomes paler / colourless; [5]
Allow ECF for -1.1 V, all answers must be consistent with the error.
- (b) no (spontaneous) reaction;
 appropriate use of Table 15 / $E^\ominus = -0.34 + 0.00 = -0.34 \text{ V}$ / E^\ominus value for the reaction would
 be negative; [2]
- (c) $\text{O}_2 / \text{Cr}_2\text{O}_7^{2-}$;
Accept names.
 E^\ominus value for the reaction with Br^- is positive / suitable calculation to show this;
 E^\ominus value for the reaction with Cl^- is negative / Cl_2 stronger oxidizing agent than
 $\text{O}_2 / \text{Cr}_2\text{O}_7^{2-}$;
 $4\text{Br}^- + \text{O}_2 + 4\text{H}^+ \rightarrow 2\text{Br}_2 + 2\text{H}_2\text{O} / \text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{Br}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O} + 3\text{Br}_2$; [5]
Award [1] for all formulas correct, [1] if coefficients correct.
- (d) sodium at negative electrode / cathode;
 chlorine at positive electrode / anode;
Accept Na and Cl_2 but not Cl.
Award [1] if electrodes not named or correct products at wrong electrodes.
 $\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$;
 $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$;
 Na : Cl_2 in ratio 2 : 1; [5]
- (e) (i) hydrogen and chlorine;
Accept formulas.
 (ratio) 1 : 1; [2]
- (ii) hydrogen and oxygen;
Accept formulas.
 (H_2 : O_2 in ratio) 2 : 1; [2]
- (iii) $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2 /$
 $4\text{OH}^- \rightarrow 2\text{H}_2\text{O} + \text{O}_2 + 4\text{e}^- /$
 $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^- /$
 $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^- .$ [1]
Accept $2\text{Cl}^- + 2\text{H}_2\text{O} \rightarrow \text{H}_2 + \text{Cl}_2 + 2\text{OH}^-$

7. (a) (i) ionic conductor / allows movement of ions between electrolytes / completes circuit; [1]
- (ii) $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-}$ (state symbols not needed); [1]
- (iii) 298 K / 25 °C, 1 atm / $1.01 \times 10^5 \text{ Pa}$, 1 mol dm⁻³ solutions; [2]
(all 3 for [2], 2 for [1])
- (iv) $0.34 - (-0.76) = 1.10 \text{ V}$; [2]
[1] for finding correct data, [1] for answer with unit (ECF).
- (v) decreases; [2]
Cu²⁺ ions are converted to Cu metal / Cu deposited on electrode;
Allow ECF from (iv).
- (vi) Cu deposited on Zn rod / rod goes pink / brown; [2 max]
blue colour of solution → paler;
gets hotter / temperature increase / exothermic;
- (b) (i) Ti²⁺ (no ECF to explanation); [2]
Ti²⁺ has greatest tendency to lose electrons / Ti³⁺ has least tendency to gain electrons;
- (ii) $\text{Ce}^{4+}(\text{aq}) + \text{Ti}^{2+}(\text{aq}) \rightarrow \text{Ce}^{3+}(\text{aq}) + \text{Ti}^{3+}(\text{aq})$; [2]
[1] for equation, [1] for state symbols. If wrong equation is given, award [1] for state symbols.
- (iii) ΔG^{\ominus} negative; [2]
reaction spontaneous / corresponds to positive cell potential;
Positive [0], non-spontaneous [1].
- (c) (i) (aqueous) sodium hydroxide / dilute sulfuric acid / sodium sulfate; [1]
Accept correct formulas.
Any combination of K⁺ / Na⁺ / H⁺ and NO₃⁻ / SO₄²⁻. Halides not acceptable.
("water" is not a solution)

(ii)



Or similar suitable diagram.

gas collection method;

names of gases correct way round at electrodes;

2:1 volume ratio correct way round;

[3]

- (d) (i) mass increases;
copper deposited;
because X is negative and attracts Cu^{2+} ions / reduction occurs at X /
 $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$;

[3]

- (ii) increase time;
increase current;

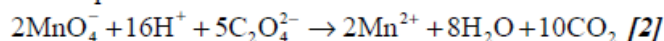
[2]

HL B 02s

6. (a) (i) (diagram) electrodes in molten salt [1];
 (diagram) polarities of electrodes and correct products shown [1];
 (diagram) electron flow in wires correctly shown [1];
 $\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$ [1];
 $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$ / $2\text{Cl}^- - 2\text{e}^- \rightarrow \text{Cl}_2$ [1]. [5]
- (ii) hydrogen [1];
 hydrogen ions are more easily discharged than sodium ions [1];
 because of position in reactivity series / OWTTE [1].
 (For last two points, accept correct reasoning based on E^\ominus values.)
Or
 oxygen [1];
 when Cl^- concentration is low [1];
 reference to difference in electrode potentials [1]. [3]
- (iii) electrons $n = \frac{5 \times 60 \times 60}{96480} = 0.187$ mol;
 (First two marks for method, $5 \times 60 \times 60$ [1] and 96480 [1])
 mass of sodium = $(0.187 \times 22.99) = 4.29 - 4.30$ g (accept 4 g) [1]; [3]
- (b) (i) $\text{C}_2\text{O}_4^{2-} \rightarrow 2\text{CO}_2 + 2\text{e}^-$ (do not accept C_2O_4) [1];
 (in ethanedioate) +3 [1];
 (in carbon dioxide) +4 [1];
 (+ sign needed for oxidation number, if omitted penalise once. If 3+ and 4+ stated, award 1 mark out of 2.)
 carbon's oxidation number increases / becomes more positive / electrons are lost [1]. [4]
- (ii) $\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 3\text{C}_2\text{O}_4^{2-} \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O} + 6\text{CO}_2$
 all formulas correct [1];
 correctly balanced [1]. [2]
- (c) (i) temperature 25°C / 298 K [1];
 concentration 1 mol dm^{-3} [1]. [2]
- (ii) (Any two of the following. Accept appropriate diagram.)
 salt bridge [1];
 containing e.g. aqueous potassium nitrate/chloride [1].
 external circuit / voltmeter [1]; [2]
- (iii) $E^\ominus = 0.80 - (-0.23) = 1.03$ V (+ sign, - sign, or no sign is acceptable) [1];
 $\text{Ni(s)} / \text{Ni}^{2+}(\text{aq}) \parallel \text{Ag}^+(\text{aq}) / \text{Ag(s)}$ species in correct order [1];
 (or reversed)
 all state symbols correct [1].
 $\text{Ni} + 2\text{Ag}^+ \rightarrow \text{Ni}^{2+} + 2\text{Ag}$ [1] [4]

7. (a) Oxidation number of Mn on left-hand side 7 [1]
 Oxidation number of C on left-hand side 3 [1]
 Oxidation number of Mn **and** C on right-hand side 2 and 4 respectively [1]

Final equation:



(If ratio $\text{MnO}_4^- : \text{C}_2\text{O}_4^{2-}$ is 2 : 5 but equation wrong give [1]. Do not penalise absence of state symbols) [5 max]

- (b) (i) Diagram must show:
 Voltmeter and metals [1]
 Salt bridge correctly labelled [1]
 $[\text{Ni}^{2+}]$ **and** $[\text{Cd}^{2+}]$ as 1 mol dm^{-3} [1] [3 max]

- (i) • H_2 gas
 • At 1 atm / 101.3 kPa
 • 298 K
 • $1 \text{ mol dm}^{-3} \text{ H}^+$
 (Four correct [2], 2/3 correct [1])
 Pt electrode [1] [3 max]

- (c) (i) $\text{Ni}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Ni}(\text{s})$ [1]
 $\text{Cd}(\text{s}) - 2\text{e}^- \rightarrow \text{Cd}^{2+}(\text{aq}) / \text{Cd}(\text{s}) \rightarrow \text{Cd}^{2+}(\text{aq}) + 2\text{e}^-$ [1]
 Allow $\text{Cd} + \text{Ni}^{2+} \rightarrow \text{Cd}^{2+} + \text{Ni}$ for [2]
 Cd is oxidised [1]
 Oxidising agent is Ni^{2+} (do not accept Ni) [1] [4 max]
- (ii) Cd is A [1]
 External electron flow from Cd to Ni [1]. [2 max]
- (iii) Overall voltage is 0.2 volts (accept + or - 0.2 V) [1]
 ΔG is negative (consequential on the sign of the potential) [1] [2 max]

- (d) Silver at cathode / - ve [1]
 Oxygen at anode / + ve [1] [2 max]

- (e) Number of coulombs = $20 \times 5 \times 60^2 / 360000$ [1]
 Number of Faradays = 3.73 [1]
 Moles of OH^- = 3.73 [1]
 Mass of NaOH = $3.73 \times 40 = 149.2 \text{ g}$ [1] [4 max]

HL B 00w

5

- (e) Cl_2 reacts with (oxidises) Br^- and I^- to the corresponding halogens [1]
 as it is a better oxidising agent than Br_2 and I_2 . (Accept balanced equations) [1]
 Br_2 reacts with I^- but not Cl^- (to form I_2) (accept balanced equation) [1]
 as it is a better oxidising agent than I_2 . [1]
 I_2 will not react with Cl^- or Br^- (as it is the weakest oxidising agent of the three). [1]

HL B .99w

Mark scheme unavailable