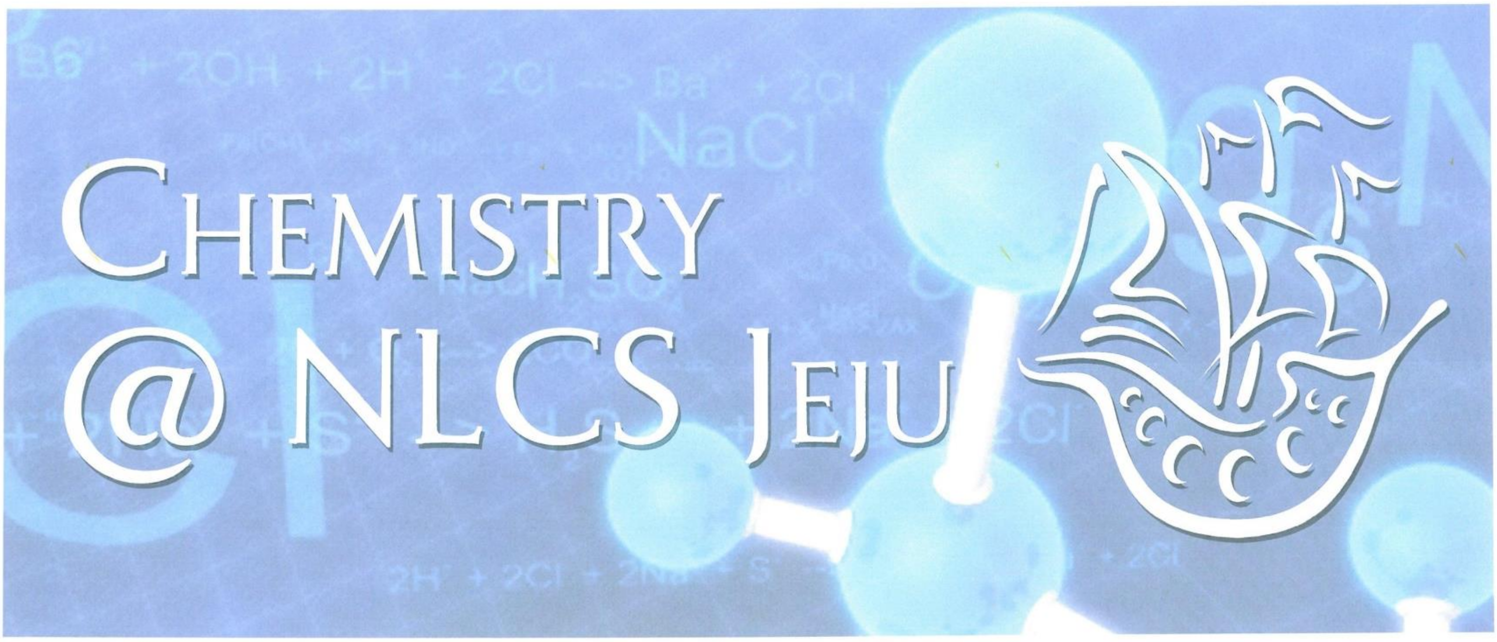
# Topic 9

# STANDARD 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 | 55/810 | **7%** | /120 | % |
| B | 137/1620 | **8%** | /240 | % |
| TOTAL | 192/2430 | **8%** | /360 | % |

Total number of papers represented here is 27, each with 30 marks of Section A and 60 marks of section B (3 questions from which you chose to answer only 1)

SL SECTION A 13s

**3.** Both sodium and sodium chloride can conduct electricity.

(a) Compare how electric current passes through sodium and sodium chloride by completing the table below. *[3]*

| | **Sodium** | **Sodium chloride** |
|---|---|---|
| **State of matter** | ........................ ........................ | ........................ ........................ |
| **Particles that conduct the current** | ........................ ........................ | ........................ ........................ |
| **Reaction occurring** | ........................ ........................ | ........................ ........................ |

(b) Sodium can be obtained by electrolysis from molten sodium chloride. Describe, using a diagram, the essential components of this electrolytic cell. *[3]*

(c) State **one** example that shows the economic importance of electrolysis. *[1]*

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SL SECTION A 13s

**4.** Ethanedioic acid (oxalic acid), $(COOH)_2$, reacts with acidified potassium permanganate solution, $KMnO_4$, according to the following equation.

$$5(COOH)_2(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 10CO_2(g) + 2Mn^{2+}(aq) + 8H_2O(l)$$

The reaction is a redox reaction.

(a) Define *oxidation* in terms of electron transfer. *[1]*

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(b) Calculate the change in oxidation numbers of carbon and manganese. *[2]*

Carbon:

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Manganese:

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(c) Identify the oxidizing and reducing agents. *[1]*

Oxidizing agent:

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Reducing agent:

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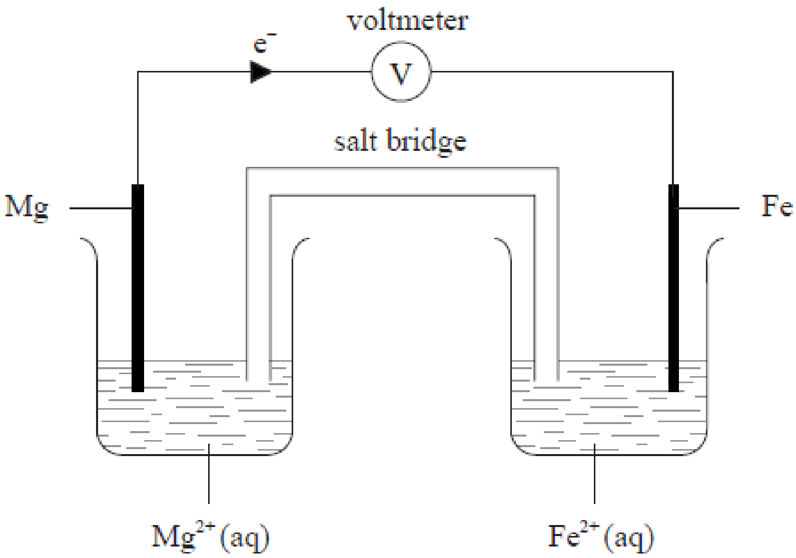
SL SECTION A 12s

4. Chemical energy can be converted to electrical energy in the voltaic cell below.




(a) (i) State the electron arrangement of a magnesium atom. *[1]*

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(ii) State the half-equation which describes the change at the Mg electrode and deduce which metal is the positive electrode (cathode) of the cell. *[2]*

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(b) Deduce the equation for the overall reaction occurring in the cell. *[1]*

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SL SECTION A 11w

**3.** (a) Molten sodium chloride can be electrolysed using graphite electrodes.

(i) Draw the essential components of this electrolytic cell and identify the products that form at each electrode. *[2]*

Product formed at positive electrode (anode):

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

Product formed at negative electrode (cathode):

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(ii) State the half-equations for the oxidation and reduction processes and deduce the overall cell reaction, including state symbols. *[2]*

Oxidation half-equation:

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Reduction half-equation:

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Overall cell reaction:

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(b) Explain why solid sodium chloride does not conduct electricity. *[1]*

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(c) Using another electrolysis reaction, aluminium can be extracted from its ore, bauxite, which contains $Al_2O_3$. State **one** reason why aluminium is often used instead of iron in many engineering applications. *[1]*

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SL 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.

Step 1: $Cu(s) + 2HNO_3(aq) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$

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

Step 3: $I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$

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

Copper:

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

Nitrogen:

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

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

Mass of brass $= 0.456\ g \pm 0.001\ g$

| **Titre** | **1** | **2** | **3** |
|---|---|---|---|
| **Initial volume of 0.100 mol dm$^{-3}$ $S_2O_3^{2-}$ ($\pm$0.05 cm$^3$)** | 0.00 | 0.00 | 0.00 |
| **Final volume of 0.100 mol dm$^{-3}$ $S_2O_3^{2-}$ ($\pm$0.05 cm$^3$)** | 28.50 | 28.60 | 28.40 |
| **Volume added of 0.100 mol dm$^{-3}$ $S_2O_3^{2-}$ ($\pm$0.10 cm$^3$)** | 28.50 | 28.60 | 28.40 |
| **Average volume added of 0.100 mol dm$^{-3}$ $S_2O_3^{2-}$ ($\pm$0.10 cm$^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) With reference to its metallic structure, describe how brass conducts electricity. *[1]*

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SL SECTION A 09w

**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** | $3.682\times10^{-1}$ |
| **Concentration of $KMnO_4$ solution / $mol\,dm^{-3}$** | $2.152\times10^{-2}$ |

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

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

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

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(d) Calculate the amount, in moles, of Fe present in the $3.682 \times 10^{-1}$ g sample of iron ore. *[2]*

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(e) Determine the percentage by mass of Fe present in the $3.682 \times 10^{-1}$ g sample of iron ore. *[2]*

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

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SL SECTION A 09sQ3

(b) Molten sodium oxide is a good conductor of electricity. State the half-equation for the reaction occurring at the positive electrode during the electrolysis of molten sodium oxide. *[1]*

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SL SECTION A 09s

**4.** (a) Define oxidation in terms of electron transfer. *[1]*

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(b) Chlorine can be made by reacting concentrated hydrochloric acid with potassium manganate(VII), $KMnO_4$.

$$2KMnO_4(aq) + 16HCl(aq) \rightarrow 2MnCl_2(aq) + 2KCl(aq) + 5Cl_2(aq) + 8H_2O(aq)$$

(i) State the oxidation number of manganese in $KMnO_4$ and in $MnCl_2$. *[2]*

$KMnO_4$ . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

$MnCl_2$ . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

(ii) Deduce which species has been oxidized in this reaction and state the change in oxidation number that it has undergone. *[2]*

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

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SL SECTION A 06s

**4.** (a) In terms of electron transfer define:

(i) *oxidation* *[1]*

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(ii) *oxidizing agent* *[1]*

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

$$Cr_2O_7^{2-} + 14H^+ + 6Fe^{2+} \rightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$$

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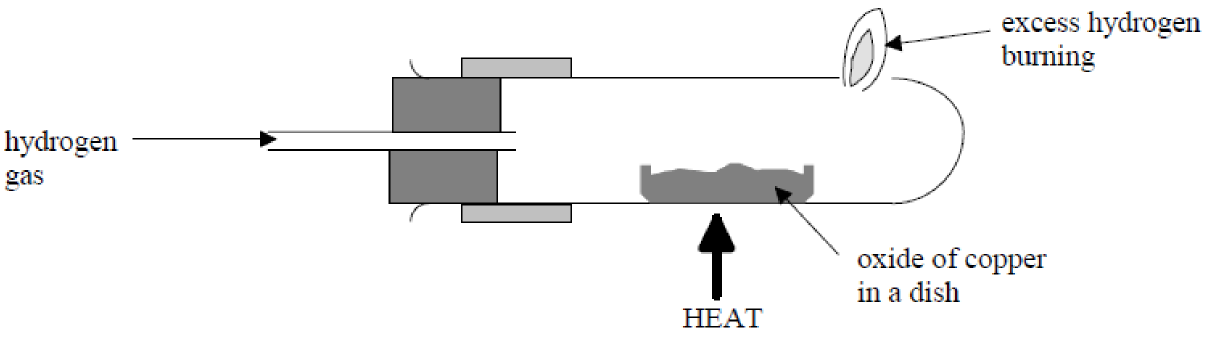
SL SECTION A 04w

**1.** An oxide of copper was reduced in a stream of hydrogen as shown below.




After heating, the stream of hydrogen gas was maintained until the apparatus had cooled.

The following results were obtained.

Mass of empty dish = 13.80 g
Mass of dish and contents before heating = 21.75 g
Mass of dish and contents after heating and leaving to cool = 20.15 g

(a) Explain why the stream of hydrogen gas was maintained until the apparatus cooled. *[1]*

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(b) Calculate the empirical formula of the oxide of copper using the data above, assuming complete reduction of the oxide. *[3]*

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(c) Write an equation for the reaction that occurred. *[1]*

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(d) State **two** changes that would be observed inside the tube as it was heated. *[2]*

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SL SECTION A 01sQ3

(c) State, giving a reason, which reactant in the following equation is acting as an oxidising agent: [2]

$$Ca + F_2 \rightarrow CaF_2$$

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

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SL SECTION A 00w

**2.** For the following reaction:

$$2Cu^{+} \rightarrow Cu + Cu^{2+}$$

(a) state the oxidation number of each species. [1]

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

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(b) write a balanced half-reaction for the oxidation process. [1]

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

(c) write a balanced half-reaction for the reduction process. [1]

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

SL B 12w

**5.** 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:

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(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 $H_3AsO_4$.

$$As_2O_3(s) + NO_3^-(aq) \rightarrow H_3AsO_4(aq) + N_2O_3(aq)$$

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

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

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SL B 12sQ6

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

$$Br_2(aq) + H_2O(l) \rightleftharpoons HOBr(aq) + HBr(aq)$$

(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.

$$2F_2(g) + 2H_2O(l) \rightarrow 4HF(g) + O_2(g)$$

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

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

(ii) 100 $cm^3$ of fluorine gas is added to water. Calculate the volume of oxygen produced at the same temperature and pressure. *[1]*

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

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SL B 11w

**5.** (a) Deduce the balanced chemical equation for the reaction between sodium and sulfur. State the electron arrangements of the reactants and product, and explain whether sulfur is oxidized or reduced. *[4]*

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

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(d) Consider the following reaction sequence:

$$\underset{\mathbf{A}}{Cr_2O_3(s)} \rightarrow \underset{\mathbf{B}}{Na_2CrO_4(aq)} \rightarrow \underset{\mathbf{C}}{Cr_2O_7^{2-}(aq)}$$

(i) State the name of **A**. *[1]*

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

(ii) Describe the redox behaviour of chromium with reference to oxidation numbers in the conversion of **B** to **C**. *[1]*

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

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(iii) Define the term *oxidizing agent* and identify the oxidizing agent in the following reaction.

$$Cr_2O_7^{2-}(aq) + I^-(aq) + 8H^+(aq) \rightarrow 2Cr^{3+}(aq) + IO_3^-(aq) + 4H_2O(l)$$ *[2]*

. . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . . .

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SL B 10wQ6

(b) Fertilizers may cause health problems for babies because nitrates can change into nitrites in water used for drinking.

(i) Define *oxidation* in terms of oxidation numbers. *[1]*

(ii) Deduce the oxidation states of nitrogen in the nitrate, $NO_3^-$, and nitrite, $NO_2^-$, ions. *[1]*

(g) Nitric acid reacts with silver in a redox reaction.

$$\_\_ Ag(s) + \_\_ NO_3^-(aq) + \_\_\_\_ \rightarrow \_\_ Ag^+(aq) + \_\_ NO(g) + \_\_\_\_$$

Using oxidation numbers, deduce the complete balanced equation for the reaction showing all the reactants and products. *[3]*

SL B 10s

**5.** (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]*

(b) Consider the following three redox reactions.

$$Cd(s) + Ni^{2+}(aq) \rightarrow Cd^{2+}(aq) + Ni(s)$$

$$Ni(s) + 2Ag^+(aq) \rightarrow Ni^{2+}(aq) + 2Ag(s)$$

$$Zn(s) + Cd^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cd(s)$$

(i) Deduce the order of reactivity of the four metals, cadmium, nickel, silver and zinc and list in order of **decreasing** reactivity. *[2]*

(ii) Identify the best oxidizing agent and the best reducing agent. *[2]*

(c) (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]*

SL B 08w

5. (a) The reaction between chlorine and bromide ions is a redox reaction.

$$Cl_2(g) + 2Br^-(aq) \rightarrow Br_2(aq) + 2Cl^-(aq)$$

Define the term *oxidation* in terms of electron transfer and identify the species that is oxidized in this reaction. *[2]*

(b) The oxidation number of oxygen is –2 in most compounds containing oxygen. Identify the oxidation numbers of all the other elements in both reactants and products in the following equation.

$$TiO_2(s) + 2Cl_2(g) + C(s) \rightarrow TiCl_4(l) + CO_2(g)$$ *[3]*

(c) By referring to oxidation numbers, deduce what happens, if anything, in terms of oxidation and reduction, to the named element in each of these reactions.

(i) Chromium in
$2K_2CrO_4(aq) + 2HCl(aq) \rightarrow K_2Cr_2O_7(aq) + 2KCl(aq) + H_2O(l)$ *[2]*

(ii) Chlorine in
$Cl_2(g) + H_2O(l) \rightarrow HCl(aq) + HClO(aq)$ *[2]*

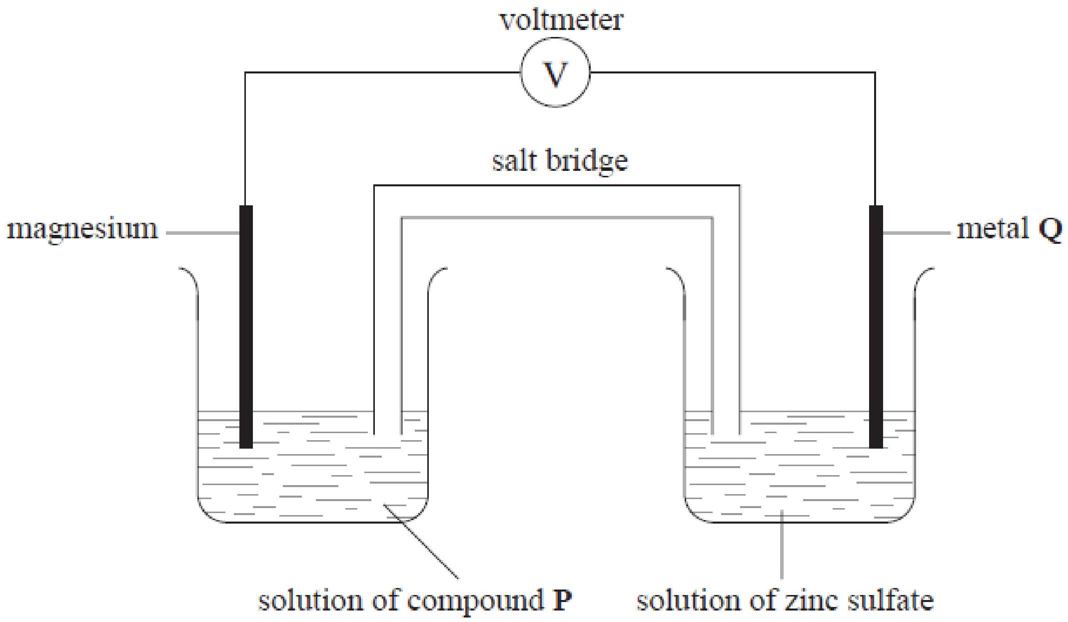
(d) The table shows some reactions involving the metals **W**, **X**, **Y** and **Z**.

| **Reaction** | **Reactants** | **Products** |
|---|---|---|
| 1 | $W + Z(NO_3)_2$ | $Z + W(NO_3)_2$ |
| 2 | $X + YCl_2$ | no reaction |
| 3 | $Y + ZSO_4$ | no reaction |
| 4 | $Z + XO$ | $X + ZO$ |

(i) Use the information to arrange the four metals in a reactivity series, starting with the most reactive. Explain with reference to each of the metals how you decided which metal was the least reactive. *[4]*

(ii) Metal **V** forms compounds in which it has an oxidation number of +3. It is more reactive than any of the metals in the table. Predict the equation for the reaction between metal **V** and the oxide of metal **X**. *[1]*

(e) The diagram shows two half-cells connected together by a salt bridge. The metal in the left-hand cell is more reactive than the metal in the right-hand cell. The reading on the voltmeter is 1.6 V.




(i) Describe the purpose of the salt bridge and identify **one** substance that might be used in it. *[2]*

(ii) Identify compound **P** and metal **Q**. *[2]*

(iii) Deduce the half-equation for the reaction in the left-hand cell. *[1]*

(iv) The voltmeter is replaced by a battery with a voltage of 2.0 V so that the reaction in part (e) (iii) is reversed. Deduce the half-equation for the reaction in the right-hand cell when the battery is connected. *[1]*

SL B 08s

**8.** (a) Consider the following reactions.

$$Ni(s) + Cu^{2+}(aq) \rightarrow Ni^{2+}(aq) + Cu(s)$$
$$Mg(s) + Zn^{2+}(aq) \rightarrow Mg^{2+}(aq) + Zn(s)$$
$$Zn(s) + Ni^{2+}(aq) \rightarrow Zn^{2+}(aq) + Ni(s)$$

(i) List the four metals in order of decreasing reactivity. *[2]*

(ii) State and explain which is the strongest reducing agent in these reactions. *[2]*

(iii) State and explain which is the strongest oxidizing agent in these reactions. *[2]*

(b) Electrolysis of molten lead(II) bromide can be carried out using platinum electrodes.

(i) Explain why lead(II) bromide does not conduct electricity in the solid state but does in the molten state. *[2]*

(ii) State a half-equation for the reaction occurring at the positive electrode (anode) and identify whether the change is oxidation or reduction. *[2]*

(iii) State a half-equation for the reaction occurring at the negative electrode (cathode) and identify whether the change is oxidation or reduction. *[2]*

SL B 06wQ4

(b)

(i) The reaction between chlorine and sodium iodide.

(ii) The reaction between silver ions and chloride ions.

(c) Deduce whether or not each of the reactions in (b) is a redox reaction, giving a reason in each case. *[4]*

(d) (i) Draw a diagram of apparatus that could be used to electrolyse molten potassium bromide. Label the diagram to show the polarity of each electrode and the product formed. *[3]*

(ii) Describe the **two** different ways in which electricity is conducted in the apparatus. *[2]*

(iii) Write an equation to show the formation of the product at each electrode. *[2]*

SL B 04s

**6.** Consider the following redox equation.

$$5Fe^{2+}(aq) + MnO_4^-(aq) + 8H^+(aq) \rightarrow 5Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(l)$$

(a) (i) Determine the oxidation numbers for Fe and Mn in the reactants and in the products. *[2]*

(ii) Based on your answer to (i), deduce which substance is oxidized. *[1]*

(iii) The compounds $CH_3OH$ and $CH_2O$ contain carbon atoms in different oxidation states. Deduce the oxidation states and state the kind of chemical change needed to make $CH_2O$ from $CH_3OH$. *[3]*

(b) A part of the reactivity series of metals, in order of decreasing reactivity, is shown below.

magnesium
zinc
iron
lead
copper
silver

If a piece of copper metal were placed in separate solutions of silver nitrate and zinc nitrate

(i) determine which solution would undergo reaction. *[1]*

(ii) identify the type of chemical change taking place in the copper and write the half-equation for this change. *[2]*

(iii) state, giving a reason, what visible change would take place in the solutions. *[2]*

(c) (i) Solid sodium chloride does not conduct electricity but molten sodium chloride does. Explain this difference, and outline what happens in an electrolytic cell during the electrolysis of molten sodium chloride using carbon electrodes. *[4]*

(ii) State the products formed and give equations showing the reactions at each electrode. *[4]*

(iii) State what practical use is made of this process. *[1]*

SL B 00s

4. (a) Use these equations, which refer to aqueous solutions, to answer the questions which follow:

$$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$$

$$Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$$

$$Mg(s) + Fe^{2+}(aq) \rightarrow Mg^{2+}(aq) + Fe(s)$$

(i) List the four metals above in order of **decreasing** reactivity. *[1]*

(ii) Define *oxidation*, in electronic terms, using **one** example from above. *[2]*

(iii) Define *reduction*, in terms of oxidation number, using **one** example from above. *[2]*

(iv) State and explain which is the **strongest reducing agent** in the examples above. *[2]*

(v) State and explain which is the **strongest oxidising agent** in the examples above. *[2]*

(vi) Deduce whether a silver coin will react with aqueous magnesium chloride. *[2]*

(b) Sketch a diagram of a cell used to electrolyse a molten salt. Label the essential components. *[4]*

(c) Describe how electrode reactions occur in an electrolytic cell and state the products at each electrode when molten copper(II) chloride is electrolysed. *[5]*

SL B .99w

**5. (a) Chlorine can be prepared in the laboratory by reacting chloride ions with an acidified solution of manganate(VII) (permanganate) ions. The unbalanced equation for the reaction is:**

$$Cl^{-}(aq) + MnO_4^{-}(aq) + H^{+}(aq) \rightarrow Mn^{2+}(aq) + Cl_2(g) + H_2O(l)$$

**Give the oxidation numbers of chlorine and manganese in the reactants and products. Write the balanced equation and identify the reducing agent in the reaction.** *[5]*

**(b) Bromine can be obtained from molten sodium bromide using electricity. Draw a clearly labelled diagram of the apparatus showing where the bromine is formed. Give the half-equation for the reaction occurring at the negative electrode (cathode) and state, with an explanation, whether this is an oxidation or reduction process. State why the products formed at the two electrodes must be kept separate from each other.** *[5]*

**(c) You are provided with aqueous solutions of chlorine, bromine and iodine and also aqueous solutions of sodium chloride, sodium bromide and sodium iodide. Describe how you could confirm experimentally the oxidising ability of the halogens and state what would be observed for each test carried out. Write equations for any reactions that occur and list the halogens in order of decreasing oxidising ability.** *[10]*

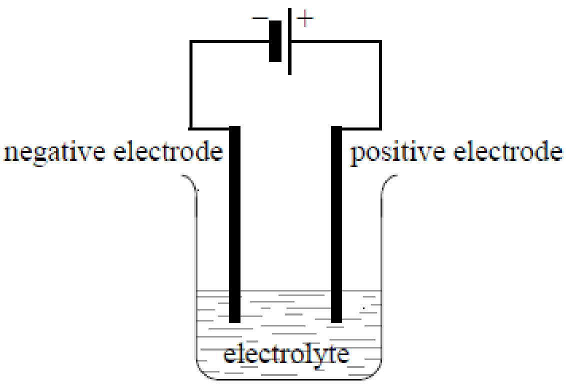
**3.** (a)

| | **Sodium** | **Sodium chloride** | |
|---|---|---|---|
| State of matter | solid (and liquid) | liquid / aqueous/solution | ; |
| Particles that conduct the current | electrons | Ions / $Na^+$ **and** $Cl^-$ | ; |
| Reaction occurring | no reaction occurs | (redox) reaction occurs / electrolysis | ; |

*[3]*

*Award [1] for each feature that is correct for both sodium **and** sodium chloride.*
*Accept equation or half-equations for the reaction of sodium chloride in "reaction occurring".*

(b)




clear diagram containing all elements (power supply, connecting wires, electrodes, container and electrolyte);
labelled positive electrode/anode **and** negative electrode/cathode;
*Accept positive and negative by correct symbols near power supply.*
*Accept power supply if shown as conventional long/short lines (as in diagram above) or clearly labelled DC power supply.*

labelled electrolyte/NaCl(l); *[3]*
*State of NaCl not needed.*

(c) production of aluminium/chlorine/lithium/magnesium/hydrogen/sodium hydroxide/ sodium chlorate / electroplating / purification of metals; *[1]*
*Do not allow production of sodium.*

SL SECTION A 13s

**4.** (a) loss of electrons; *[1]*

(b) *Carbon:*
III to IV / +3 to +4 / (+)1;

*Manganese:*
VII to II / +7 to +2 / –5; *[2]*
*Penalize incorrect notation such as 3+ once only.*

(c) *Oxidizing agent:* $MnO_4^-$ **and** *Reducing agent:* $(COOH)_2$; *[1]*
*Accept correct names instead of formulas.*
*Do not accept Mn and C.*

SL SECTION A 12s

**4.** (a) (i) 2,8,2; *[1]*

(ii) $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$;
Fe/iron; *[2]*
*Do not accept Fe/$Fe^{2+}$ half-equation or $Fe^{2+}$.*

(b) $Mg(s) + Fe^{2+}(aq) \rightarrow Mg^{2+}(aq) + Fe(s)$; *[1]*

SL SECTION A 11w

**3.** (a) (i) *Cell showing:*
container, liquid, electrodes **and** power supply;
*No labels are required, but do not award mark if incorrect labels are used (e.g. sodium chloride solution). A line must be drawn on the container to indicate the presence of a liquid. If power supply is a battery then do not penalize electrodes incorrectly assigned as + or –.*

*Positive electrode (anode):*
chlorine (gas) / $Cl_2(g)$

**and**

*Negative electrode (cathode):*
sodium (liquid) / $Na(l)$; *[2]*
*Ignore state symbols in (i) but do not award mark for Cl.*

(ii) *Oxidation half-equation:*
$2Cl^- \rightarrow Cl_2 + 2e^-$ / $Cl^- \rightarrow \frac{1}{2}Cl_2 + e^-$

**and**

*Reduction half-equation:*
$Na^+ + e^- \rightarrow Na$ / $2Na^+ + 2e^- \rightarrow 2Na$;
*Allow e instead of $e^-$.*

*Overall cell reaction:*
$2NaCl(l) \rightarrow 2Na(l) + Cl_2(g)$ / $NaCl(l) \rightarrow Na(l) + \frac{1}{2}Cl_2(g)$; *[2]*
*Award **[1]** for oxidation **and** reduction half-equations.*
*Award **[1]** for overall cell reaction, including correct state symbols.*
*Accept $Na^+(l)$ + $Cl^-(l)$ instead of NaCl(l) as a reactant.*
*Penalize equilibrium arrows once only.*

(b) ions not free to move when solid / ions in rigid lattice / *OWTTE*; *[1]*

(c) Al less dense (compared to Fe) / Al forms a protective (oxide) layer / Fe rusts / *OWTTE*; *[1]*
*Do not accept Al is lighter.*

**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–; *[2]*
*Penalize missing + sign or incorrect notation such as 2+, $2^+$ or II, once only.*

(ii) nitric acid/$HNO_3$ / $NO_3^-$/nitrate; *[1]*
*Allow nitrogen from nitric acid/nitrate but not just nitrogen.*

(b) (i) $0.100 \times 0.0285$;
$2.85 \times 10^{-3}$ (mol); *[2]*
*Award [2] for correct final answer.*

(ii) $2.85 \times 10^{-3}$ (mol); *[1]*

(iii) $(63.55 \times 2.85 \times 10^{-3}) = 0.181$ 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) *Brass has:*
delocalized electrons / sea of mobile electrons / sea of electrons free to move; *[1]*
*No mark for just "mobile electrons".*

SL SECTION A 09w

**1.** (a) $MnO_4^-(aq) + 5Fe^{2+}(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)$ *[2]*

*Award **[2]** if correctly balanced.*
*Award **[1]** for correctly placing $H^+$ and $H_2O$.*
*Award **[1 max]** for correct balanced equation but with electrons shown.*
*Ignore state symbols.*

(b) $Fe^{2+}$ / iron(II); *[1]*
*Do not accept iron.*

(c) $n = 2.152 \times 10^{-2} \times 2.250 \times 10^{-2}$;
$4.842 \times 10^{-4}$ (mol); *[2]*
*Award **[1]** for correct volume*
*Award **[1]** for correct calculation.*

(d) 1 mol of $MnO_4^-$ reacts with 5 mol of $Fe^{2+}$;
$5 \times 4.842 \times 10^{-4} = 2.421 \times 10^{-3}$ (mol); *[2]*
(same number of moles of Fe in the iron ore)
*Allow ECF from part (a) and (c) provided some mention of mole ratio is stated.*

(e) $2.421 \times 10^{-3} \times 55.85 = 0.1352$ (g);
$\frac{0.1352}{0.3682} \times 100 = 36.72\%$; *[2]*
*Allow ECF from part (d).*

SL SECTION A 09sQ3

(b) $2O^{2-} \rightarrow O_2 + 4e^-$ / $O^{2-} \rightarrow \frac{1}{2}O_2 + 2e^-$ ; *[1]*
*Accept e instead of $e^-$.*

SL SECTION A 09s

**4.** (a) Loss of (one or more) electrons; *[1]*

(b) (i) ($KMnO_4$) +7;
($MnCl_2$) +2; *[2]*
*Must have + sign for mark.*
***[1 max]** if roman numerals or 7+ or 2+ used or if + signs are missing.*

(ii) $Cl^-$ / chloride / chlorine / Cl (has been oxidized) / HCl ;
oxidation number from –1 to 0 / has <u>increased</u> by one ; *[2]*
*If HCl is given for first mark, it must be clear that it is the Cl that has the change of oxidation number.*

SL SECTION A 06s

**4.** (a) (i) loss of electrons; *[1]*

(ii) (a species that) gains electrons (from another species) / causes electron loss; *[1]*

(b) changes by 3;
reduced because its oxidation number decreased / $+6 \rightarrow +3$ / $6+ \rightarrow 3+$ / it has gained electrons; *[2]*

SL SECTION A 04w

**1.** (a) to prevent (re)oxidation of the copper / *OWTTE*; ***[1]***

(b) number of moles of oxygen $= \frac{1.60}{16.00} = 0.10$;

number of moles of copper $= \frac{6.35}{63.55} = 0.10$;

empirical formula $= \text{Cu}(0.10) : \text{O}(0.10) = \text{CuO}$; ***[3]***

*Allow ECF.*

*Award **[1]** for CuO with no working.*

*Alternate solution*

$\frac{6.35}{7.95} = 79.8\,\%$ $\quad \frac{1.60}{7.95} = 20.2\,\%$

$\frac{79.8}{63.5} = 1.25$ $\quad \frac{20.2}{16} = 1.29$

(c) $H_2 + CuO \rightarrow Cu + H_2O$; ***[1]***

*Allow ECF.*

(d) (black copper oxide) solid turns red / brown;
condensation / water vapour (on sides of test tube); ***[2]***
*Accept change colour.*
*Do **not** accept reduction of sample size.*

SL SECTION A 01sQ3

(c) Fluorine/ $F_2$ ***[1]***

$F_2$ gains electrons / $F_2$ is reduced / oxidation number decreases ***[1]***

**or**

Ca loses electrons / Ca oxidation number increases ***[1]*** ***[2 max]***

SL SECTION A 00w

**2.** (a) $Cu^+$ $+1$; Cu0, $Cu^{2+}$ $+2$. *(any two correct **[1]**, + sign needed)* ***[1]***

(b) $Cu^+ \rightarrow Cu^{2+} + e^-$ ***[1]***

(c) $Cu^+ + e^- \rightarrow Cu$ ***[1]***

**5.** (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/$e^-$ acceptor / causes oxidation / oxidation number/state decreases **and** Reducing agent: electron/e/$e^-$ 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; *[3]*
*Accept $HNO_3(aq)$/nitric acid.*
*Accept arsenic oxide.*
*Species must be fully correct to score M3.*
*Ignore state symbols.*

SL B 12sQ6

(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 ($cm^3$) / 0.050 $dm^3$; *[1]*

SL B 11w

**5.** (a) $2Na(s) + S(s) \rightarrow Na_2S(s)$ / $2Na(s) + \frac{1}{2}S_2(g) \rightarrow Na_2S(s)$ / $16Na(s) + S_8(s) \rightarrow 8Na_2S(s)$ ;
*Ignore state symbols.*

Na: 2, 8, 1 **and** S: 2, 8, 6 ;

$Na^+$: 2, 8 **and** $S^{2-}$ : 2, 8, 8 ;
reduced since it has gained electrons / reduced since oxidation number has decreased; *[4]*
*Do not award mark if incorrect oxidation numbers are given.*

(d) (i) chromium(III) oxide; *[1]*
*Do not award mark for chromium oxide.*

(ii) chromium is neither oxidized or reduced since there is no change in oxidation number/+6 to +6; *[1]*

(iii) substance reduced / causes other substance to be oxidized / increase oxidation number of another species / gains electrons / *OWTTE*;

*Oxidizing agent:*
$Cr_2O_7^{2-}$ / dichromate (ion); *[2]*

SL B 10wQ6

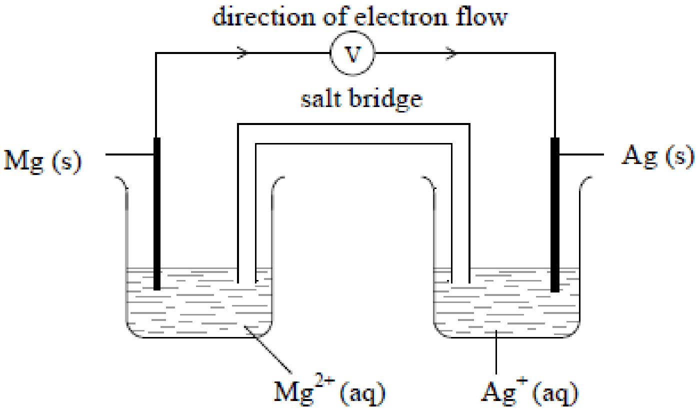
(b) (i) increase in the oxidation number; *[1]*

(ii) $(NO_3)+5$ **and** $(NO_2^-)+3$; *[1]*
*Accept V and III.*
*Do not penalize missing charges on numbers.*

(g) change in oxidation numbers: Ag from 0 to +1 **and** N from +5 to +2;
*Do not penalise missing charges on numbers.*

balanced equation: $3Ag + NO_3^- + 4H^+ \rightarrow 3Ag^+ + NO + 2H_2O$ *[3]*
*Award **[1]** for correct reactants and product;*
*Award **[3]** for correct balanced equation.*
*Ignore state symbols.*

5. (a) (i)




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

direction of electron flow; *[4]*

(ii) *Oxidation:*
$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^{-}$;

*Reduction:*
$Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$; *[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 (c) (ii).*

(b) (i) Zn > Cd > Ni > Ag
Zn most reactive;
rest of order correct; *[2]*

(ii) *Best oxidizing agent:*
$Ag^{+}$;
*Do not accept Ag.*

*Best reducing agent:*
Zn; *[2]*
*Do not accept $Zn^{2+}$.*

(c) (i) sodium chloride crystals consist of <u>ions</u> in a (rigid) lattice / <u>ions</u> cannot move (to electrodes) / *OWTTE*;
when melted <u>ions</u> free to move / <u>ions</u> move when potential difference/voltage applied; ***[2]***

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

electrons 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)*:
$2Cl^- \rightarrow Cl_2 + 2e^-$ / $Cl^- \rightarrow \frac{1}{2}Cl_2 + e^-$;

*(Negative electrode/cathode):*
$2Na^+ + 2e^- \rightarrow 2Na$ / $Na^+ + e^- \rightarrow Na$; ***[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 (c) (ii).*

**5.** (a) (oxidation is) loss of electron(s);
$Br^-$ / bromide; ***[2]***

(b) Ti +4 and +4;
Cl 0 and –1;
C 0 and +4; ***[3]***
*Penalize missing +, or answers written as 4+ once only.*
*If no marks scored allow **[1]** if all oxidation numbers for reactants or for products are correct.*

(c) (i) Cr oxidation number +6/same on both sides/does not change;
neither oxidation nor reduction occur; ***[2]***
*No ECF.*

(ii) Cl oxidation number 0 on left and –1 and +1 on right;
both oxidation and reduction occur / disproportionation; ***[2]***
*No ECF.*

(d) (i) $W > Z > Y > X$;
*Award **[1]** mark for correct order.*

X below Y because of reaction 2/because X will not displace Y;
X below Z because of reaction 4/because X displaced by Z;
X below W because of reaction 1/because Z displaced by W and Z displaces X;
Y below Z because of reaction 3/because Y will not displace Z; ***[4 max]***
*Give credit for OWTTE in this part.*
*Any three of last four score **[1]** each.*

(ii) $2V + 3XO \rightarrow V_2O_3 + 3X$; ***[1]***

(e) (i) allows ions to flow through it / to complete the circuit/balances charge;
potassium chloride/KCl / potassium nitrate /$KNO_3$; ***[2]***
*Accept any other unreactive soluble salt including sulfates.*

(ii) P is magnesium sulfate/$MgSO_4$;
*Accept magnesium chloride $MgCl_2$ or magnesium nitrate $Mg(NO_3)_2$.*

Q is zinc/Zn; ***[2]***

(iii) $Mg \rightarrow Mg^{2+} + 2e^-$; ***[1]***
*Accept e instead of $e^-$.*
*Accept $-2e^-$ on left.*
*Ignore state symbols.*

(iv) $Zn \rightarrow Zn^{2+} + 2e^-$; ***[1]***
*Accept e instead of $e^-$.*
*Accept $-2e^-$ on left.*
*Ignore state symbols.*

SL B 08s

**8.** (a) (i) Mg > Zn > Ni > Cu *[2]*

*Four metals in correct order, award* ***[2]****, first and last metal order correct award* ***[1]****.*

(ii) Mg;

Mg can reduce all other species/has a greater tendency to donate electrons; *[2]*

(iii) $Cu^{2+}$;

$Cu^{2+}$ can oxidise other species/has a greater tendency to accept electrons; *[2]*

*Do not accept Cu*

(b) (i) (solid state) ions in fixed position;

(molten state) ions are free to move; *[2]*

(ii) $2Br^- \rightarrow Br_2 + 2e^-$ / $Br^- \rightarrow \frac{1}{2}Br_2 + e^-$;

*Accept e instead of $e^-$*

oxidation; *[2]*

(iii) $Pb^{2+} + 2e^- \rightarrow Pb$;

*Accept e instead of $e^-$*

reduction; *[2]*

*Award* ***[1]*** *for two correct equations with wrong electrodes.*

SL B 06w

(c) reaction in (i) is redox;

chlorine is reduced/gains electrons/decreases its oxidation number /

(sodium) iodide is oxidized/loses electrons/increases its oxidation number;

reaction in (ii) is not redox;

no electron transfer/change in oxidation number; *[4]*

(d) (i) (diagram showing)

container, liquid, electrodes and power supply;

bromine formed at + electrode;

potassium formed at – electrode; *[3]*

*Award* ***[1]*** *for both correct products shown at wrong electrodes, or if no polarity indicated.*

(ii) electrons flow through connecting wires;

ions move (through liquid) to electrodes (and lose/gain electrons); *[2]*

(iii) $K^+ + e^- \rightarrow K$;

$2Br^- \rightarrow Br_2 + 2e^-$; *[2]*

*No need to indicate polarity of electrodes.*

*Accept e instead of $e^-$.*

**6.** (a) (i) Fe reactant +2 AND Fe product +3 AND Mn product +2;
Mn reactant +7; *[2]*
*Do not accept Roman numerals.*

(ii) $Fe^{2+}$ / iron(II) ions / ferrous ions; *[1]*
*Do not accept "iron".*

(iii) $CH_3OH$ oxidation state −2;
$CH_2O$ oxidation state 0;
(change is) oxidation / dehydrogenation; *[3]*

(b) (i) silver nitrate; *[1]*

(ii) oxidation;
$Cu \rightarrow Cu^{2+} + 2e$; *[2]*

(iii) (silver nitrate) solution turns blue / grey or black or silver solid forms;
copper ions form / $Cu^{2+}$ ions form / silver deposited; *[2]*

(c) (i) sodium chloride crystals consist of ions in a rigid lattice / ions can not move about;
when melted the ions are free to move or ions move when a voltage is applied;
in electrolysis positive sodium ions or $Na^+$ ions move to the negative electrode or cathode;
and negative chloride ions or $Cl^-$ move to the positive electrode or anode; *[4]*

(ii) sodium formed at cathode or negative electrode;
$Na^+ + e \rightarrow Na$;
chlorine formed at anode or positive electrode;
$2Cl^- \rightarrow Cl_2 + 2e$; *[4]*
*1st and 3rd marks can be scored in (c) (i).*

(iii) manufacture of sodium and chlorine / one stated use of chlorine or sodium; *[1]*

**4.** (a) (i) Mg, Fe, Cu, Ag *[1]*

(ii) Loss/removal of electrons; *[1]*
Fe/Cu/Mg *[1]*

(iii) Reduction/decrease in oxidation number; *[1]*
$Cu^{2+}$ / $Ag^{+}$ / $Fe^{2+}$ *[1]*

(iv) Mg *[1]*

It reduces $Fe^{2+}$; Fe reduces $Cu^{2+}$; Cu reduces $Ag^{+}$. *[1]*

**OR** It reduces $Fe^{2+}$ to Fe, which can then reduce everything else. *[1]*

**OR** It is the most reactive (of the metals shown). *[1]*

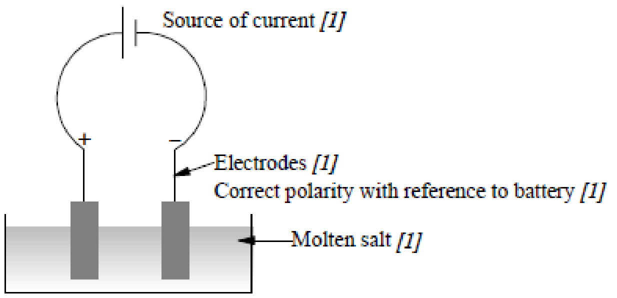
(v) $Ag^{+}$ *[1]*

Every metal present can reduce it to Ag. *[1]*

(vi) Does not react *[1]*

It has too low an electrode potential/too low in reactivity series/less reactive than Mg. *[1]*

(b)




(c) At the cathode; *[1]*
electrons are given to the cations/positive ions; *[1]*
electrons are removed from the anions/negative ions; *[1]*
(at the anode)
copper formed (at the cathode); *[1]*
chlorine formed (at the anode); *[1]*

Summary: electron loss *[1]* electron gain *[1]*
copper formed *[1]* chlorine formed *[1]*
related to correct electrode in **one** case *[1]*
(*e.g.* Cu at cathode **OR** $Cu^{2+}$ gains $2e^{-}$ at cathode.)

**5.** (a) $Cl^-(-1)$, $Cl_2(0)$ *[1]*
$MnO_4^-(+7)$, $Mn^{2+}(+2)$ *[1]*
$10Cl^-(aq) + 2MnO_4^-(aq) + 16H^+(aq) \rightarrow 2Mn^{2+}(aq) + 5Cl_2(g) + 8H_2O(l)$ *[2]*
$Cl^-$ reducing agent *[1]*

(b) Diagram should show:
External source of electric current and circuit *[1]*
Bromine evolved at + electrode (anode) *[1]*
$Na^+ + e^- \rightarrow Na$ *[1]*
Reduction because $Na^+$ has gained an electron/oxidation number is decreased. *[1]*
Because sodium and bromine would recombine/they are reactive *[1]*

(c) $Cl_2$ should be added to NaBr(aq) *[1]*
A yellow/red colour of bromine will be observed. *[1]*
$Cl_2 + 2Br^- \rightarrow Br_2 + 2Cl^-$ *[1]*

$Cl_2$ should be added to NaI(aq) *[1]*
A red/brown *(accept orange)* colour of iodine will be observed *[1]*
$Cl_2 + 2I^- \rightarrow I_2 + 2Cl^-$ *[1]*

$Br_2$ should be added to NaI(aq) *[1]*
A red/brown *(accept orange)* colour of iodine will be observed. *[1]*

*accept "darker colour formed" since colour change is sometimes difficult to see*

$Br_2 + 2I^- \rightarrow I_2 + 2Br^-$ *[1]*
$Cl_2 > Br_2 > I_2$ *[1]*