IB · DP · Chemistry

3 hours 16 questions

Structured Questions: Paper 2

# 9.1 Redox Processes

9.1.1 Oxidation & reduction / 9.1.2 Deducing Oxidation Numbers / 9.1.3 Oxidizing & Reducing agents / 9.1.4 Naming Transition Metal Compounds / 9.1.5 Half Equations / 9.1.6 The Activity Series / 9.1.7 Redox Titrations / 9.1.8 The Winkler Method / 9.1.9 Voltaic Cells / 9.1.10 Electrolytic Cells

| | |
|---|---|
| Easy (5 questions) | /45 |
| Medium (6 questions) | /73 |
| Hard (5 questions) | /54 |
| **Total Marks** | **/172** |

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# Easy Questions

**1 (a)** Chlorine reacts with water to form chlorine water via the following equation.

$$Cl_2 + H_2O \rightarrow HOCl + HCl$$

State the oxidation number of chlorine in the following species

$Cl_2$ ......................

HOCl .....................

HCl ......................

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**(3 marks)**

**(b)** Chlorine is an *oxidising agent*.

Define *oxidising agent* in terms of electrons.

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**(1 mark)**

**(c)** Nitrogen monoxide, NO, is formed when silver metal reduces nitrate ions, $NO_3^-$ , ions in an acidic solution. State the oxidation numbers of nitrogen in NO and $NO_3^-$.

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**(2 marks)**

**(d)** State the half equation for the formation of silver ions, $Ag^+$ (aq), from silver metal.

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**(1 mark)**

**2 (a)** Deduce the oxidation numbers of of the elements in the following species.

$S^{2-}$ …………………

$Sn^{2+}$ …………………

$V^{3+}$ …………………

Si …………………

$Sb^{3+}$ …………………

$H^{-}$ …………………

(6 marks)

**(b)** Oxidation states are sometimes visible in the names of chemicals. Deduce the oxidation numbers of the stated elements in the following species.

Copper in copper(I) oxide …………………

Iron in iron(III) oxide …………………

Phosphorus(V) oxide …………………

**(3 marks)**

**(c)** The dichromate(VI) ion, $Cr_2O_7^{2-}$ (aq) , reacts with sulfite ions, $SO_3^{2-}$ (aq), as follows.

$$Cr_2O_7^{2-} (aq) + 8H^+ (aq) + 3SO_3^{2-} (aq) \rightarrow 2Cr^{3+} (aq) + 4H_2O (l) + 3SO_4^{2-} (aq)$$

i) State whether the sulfite ions, $SO_3^{2-}$ (aq), are acting as an oxidising or reducing agent.

ii) Justify your answer to part (i).

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**(2 marks)**

**(d)** Redox reactions can be identified by either reduction and oxidation occurring or the presence of a reducing agent and an oxidising agent.

i) Deduce if the reaction between hydrochloric acid and sodium hydroxide is a redox reaction.

$$HCl + NaOH \rightarrow NaCl + H_2O$$

ii) Justify your answer.

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**(3 marks)**

**3 (a)** An iron tablet, weighing 1.35 g was dissolved in dilute sulfuric acid. The sample was dissolved in sulfuric acid to oxidise all of the iron to $Fe^{2+}$ ions.

The solution is then titrated with 0.02 mol $dm^{-3}$ potassium dichromate, $K_2Cr_2O_7$, producing $Fe^{3+}$ and $Cr^{3+}$ ions in acidic solution. The titration requires 31.00 $cm^3$ of $K_2Cr_2O_7$ for 1.35 g of the sample.

Balance the following half equations:

$Fe^{2+}$ (aq) → $Fe^{3+}$ (aq) + .........

$Cr_2O_7^{2-}$(aq) + 14$H^+$ (aq) + ......... → .......$Cr^{3+}$ (aq) + ......$H_2O$ (l)

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**(1 mark)**

**(b)** The overall equation for the reaction in part (a) is as follows.

$6Fe^{2+}$ (aq) + $Cr_2O_7^{2-}$ (aq) + 14$H^+$ (aq) → $6Fe^{3+}$ (aq) + $2Cr^{3+}$ (aq) + $7H_2O$ (l)

i) Using the information in part (a), calulate the number of moles of potassium dichromate, $K_2Cr_2O_7$ used.

ii) Use your answer to part (b) (i) to determine the number of moles of $Fe^{2+}$ in the sample.

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**(2 marks)**

**(c)** Using the information in part (a) calculate the mass, in grams, of iron in the original sample.

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**(2 marks)**

**(d)** Using the information in part (a) and your answer to part (c) calculate the percentage of iron in the original 1.35 g iron tablet.

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**(1 mark)**

**4 (a)** Zinc metal will react with copper sulfate solution. State the equation for this reaction.

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**(2 marks)**

**(b)** Predict the products, if any, of the reaction between lead(IV) oxide and zinc.

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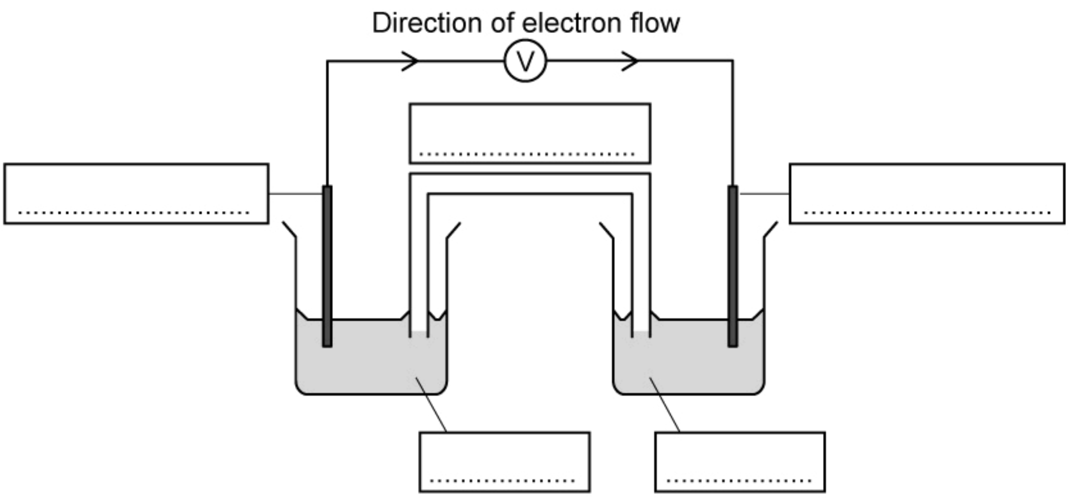
**(1 mark)**

**(c)** A voltaic cell is made from a half-cell containing a zinc electrode in a solution of zinc nitrate and a half-cell containing a silver electrode in a solution of silver nitrate. Using section 25 in the data booklet, label the following diagram.

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**(5 marks)**

**(d)** State three differences between a voltaic cell and an electrolytic cells.

**(3 marks)**

**5 (a)** A 500 $cm^3$ sample of lake water was analysed using the Winkler Method. 0.0600 mol $dm^{-3}$ sodium thiosulfate solution, $Na_2S_2O_3$ (aq), was titrated against liberated iodine. It was determined that the average titre of $Na_2S_2O_3$ was 19.30 $cm^3$.

Calculate the amount, in moles, of sodium thiosulfate used in the titration.

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**(1 mark)**

**(b)** In the Winkler method 1 mol of $O_2$ is equivalent to 4 mol of $Na_2S_2O_3$ Deduce the amount, in moles, of oxygen that has reacted.

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**(1 mark)**

**(c)** Use your answer to part (b) and section 6 in your data booklet to calculate the following

- The concentration, in mol $dm^{-3}$, of oxygen in the water ...................
- The concentration, in g $dm^{-3}$, of oxygen in the water ....................

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**(2 marks)**

**(d)** Define the *biological oxygen demand (BOD)*.

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**(3 marks)**

# Medium Questions

**1 (a)** Common household bleach is a cleaning product which smells like chlorine gas and is therefore, also called chlorine bleach.
It contains a mixture of sodium chlorate (NaO*Cl*), sodium chloride and water and can be made by dissolving chlorine gas in a solution of sodium hydroxide.

i) Write a balanced equation with state symbols for this reaction.
ii) Deduce the oxidation number of chlorine in all of the chlorine-containing reactants and products

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**(3 marks)**

**(b)** The mixing of household bleach with ammonia during cleaning should be avoided, as a redox reaction between the ammonia and the chlorate(I) ions in bleach will generate toxic chlorine gas and hydrazine ($N_2H_4$).

The overall redox reaction for this reaction is shown below.

$$2NH_3\,(aq) + 2ClO^-\,(aq) \rightarrow N_2H_4\,(aq) + Cl_2\,(g) + 2OH^-\,(aq)$$

i) What are the oxidation numbers of the nitrogen atom in $NH_3$ and in $N_2H_4$?
ii) What is the oxidizing agent in this reaction? Explain your answer.
iii) Why is the hazard of the toxic chlorine gas being produced greater than the hazard of hydrazine?

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**(4 marks)**

**(c)** Due to the risks associated with chlorine-based bleach, alternative bleaches are often used instead. These bleaches are based on peroxides such as hydrogen peroxide.

Manganate(VII) ions oxidize hydrogen peroxide to oxygen gas. The reaction is carried out with both species under acidic conditions.

i) Identify the oxidizing and reducing agents in this reaction.

ii) Write the half-equation for the oxidation of hydrogen peroxide to oxygen gas.

iii) The manganate(VII) ions themselves get reduced to manganese(II) ions. Write down the half-equation for the reduction of manganate(VII) ions.

iv) Deduce the overall redox equation for this reaction.

**(5 marks)**

**(d)** Explain how the oxidation number of the oxygen atom in $H_2O_2$ is different from its oxidation state in other compounds.

**(2 marks)**

**2 (a)** Metals can often be seen written as a list, from the most reactive metal to the least reactive metal. This list is known as the reactivity series of metals and can be used to predict the feasibility of a reaction.

Below is a section of the reactivity series of metals, ordered from most to least reactive:

Calcium

Magnesium

Aluminium

Zinc

Iron

Tin

Lead

A piece of zinc was placed into a solution of iron(II) sulfate and a solution of magnesium sulfate.

Predict, giving a reason, whether a reaction would occur in each solution.

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**(2 marks)**

**(b)** Copper is below lead on the reactivity series shown in part (a). A piece of zinc was placed into a solution of copper(II) sulfate. Write the half equation for the zinc and identify the type of reaction taking place.

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**(2 marks)**

**(c)** Many chemical reactions are redox reactions as they involve the transfer of electrons.

i) Explain the role of the oxidizing agent in a redox reaction in terms of electron transfer.

ii) State the most common oxidation number of an oxygen atom when in a compound.

iii) Which oxygen compounds are an exception to your answer in part (ii)? Explain your answer.

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**(4 marks)**

**(d)** The following reaction is an example of a common redox reaction:

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

Deduce the oxidation numbers of iron and manganese in the above reaction, both as reactants and as products.

State which substance is reduced.

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**(3 marks)**

**(e)** The amount of iron in some dietary iron supplements was analyzed by redox titration. Four tablets were crushed and dissolved in 50.0 $cm^3$ of 2.00 mol $dm^{-3}$ sulfuric acid. The solution was then transferred to a 250 $cm^3$ volumetric flask and made up to 250 $cm^3$ with distilled water.

A 25.0 $cm^3$ sample of the iron tablets solution was titrated against 0.00500 mol $dm^{-3}$ potassium manganate(VII) and 25.8 $cm^3$ was needed for complete reaction.

Determine the amount of iron, in mol, in **one** tablet.

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**(4 marks)**

**3 (a)** Halide ions, such as chloride, $Cl^-$, can be identified using chemical tests. If an unknown compound is dissolved in dilute nitric acid, and then silver nitrate solution is added, a precipitate will form if the unknown solution contains halide ions. The precipitate formed will be a silver halide.

The general equation for the precipitation reaction of halide ions with silver nitrate solution is:

$$AgNO_3\,(aq) + X^-\,(aq) \rightarrow AgX\,(s) + NO_3^-\,(aq)$$

i) Deduce the oxidation number of silver in $AgNO_3$ and AgX and deduce the oxidation number of the halide in $X^-$ and in AgX.

ii) Is the precipitation of silver halides a redox reaction? Explain your answer.

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**(4 marks)**

**(b)** Halide ions can also react with other halogens in aqueous solutions. Chlorine reacts in a redox reaction with an aqueous solution of sodium bromide, to form sodium chloride and bromine.

$$Cl_2\,(aq) + NaBr\,(aq) \rightarrow NaCl\,(aq) + Br_2\,(aq)$$

i) State what type of redox reaction this is.

ii) Using the overall redox reaction above, deduce the half-equation for chlorine. State whether chlorine is oxidized or reduced.

iii) Using the overall redox reaction above, deduce the half-equation for bromine. State whether bromine is oxidized or reduced.

iv) Use the reaction above and your knowledge of the halogens, to explain whether chlorine or bromine is a stronger oxidizing agent.

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**(7 marks)**

**(c)** Chlorine also oxidizes sulfur dioxide ($SO_2$) in aqueous solutions to sulfate ions ($SO_4^{2-}$) under acidic conditions.

i) Deduce the half-equation for the reduction of chlorine in aqueous solution.

ii) Deduce the half-equation for the oxidation of sulfur dioxide in aqueous solution.

**(2 marks)**

**(d)** Use the two half-equations from part (c) to construct the overall redox equation for this reaction.

**(1 mark)**

4 (a) The iron of railway lines rusts when it comes into contact with water and oxygen. The overall redox equation for the rusting of iron is as follows:

$$4Fe\ (s) + 3O_2\ (g) + 6H_2O\ (g) \rightarrow 4Fe(OH)_3\ (s)$$

Define the term *reduction*.

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**(1 mark)**

**(b)** State, with a reason, the oxidizing agent in this reaction in part (a).

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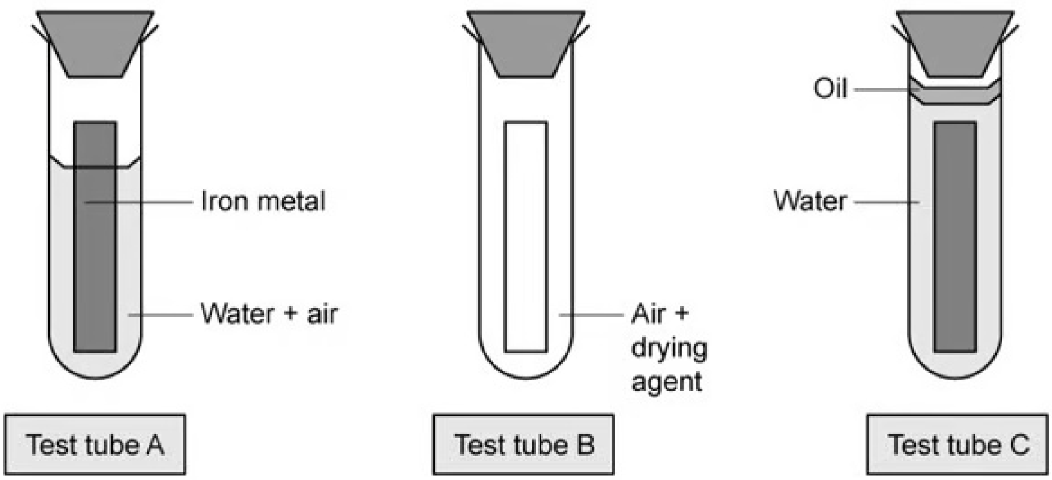
**(2 marks)**

**(c)** A student investigates the rate of rusting of a piece of iron under different conditions.

**Figure 1** shows the set-up of the students' experiment.

**Figure 1**




Predict in which test tube(s) the iron metal will not rust. Explain your answer.

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(3 marks)

(d) Deduce the oxidation number of each of the stated elements in the ions and compounds to complete **Table 1** below.

**Table 1**

| **Species** | **Oxidation number** |
|---|---|
| Oxygen in $Na_2O_2$ | |
| Hydrogen in $MgH_2$ | |
| Nitrogen in $NO_3^-$ | |
| Chlorine in *Cl*F | |

(4 marks)

**5 (a)** Aluminium is present in the Earth's crust in aluminium ore, called bauxite. A number of processes are done to this ore, to extract the aluminium from it. The bauxite is initially purified to produce aluminium oxide, $Al_2O_3$. Electrolysis is then carried out on the molten $Al_2O_3$, to extract the aluminium.

i) Write down the overall equation for the extraction of aluminium from aluminium oxide by electrolysis.

ii) State whether the aluminium oxide is oxidized or reduced in the electrolysis reaction. Explain your answer.

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**(3 marks)**

**(b)** Another ionic compound which can undergo electrolysis is molten lead bromide.

i) Explain, in terms of ions and electrons, what would happen in an electrolytic cell during the electrolysis of lead bromide, using carbon electrodes.

ii) State two different ways in which electrical charge flows in the electrolysis apparatus.

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**(4 marks)**

**(c)** State the products formed at each electrode during the electrolysis of molten lead bromide, giving the equations at each electrode with state symbols.

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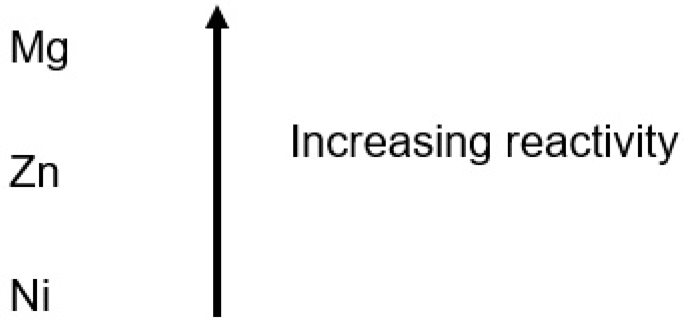
(3 marks)

**(d)** Draw a labelled diagram of the apparatus suitable to carry out the electrolysis of molten lead bromide. Include the direction of electron flow, the negative electrode (cathode), the positive electrode (anode) and the electrolyte.

(3 marks)

**6 (a)** The list below shows three metals from the activity series in order of reactivity.




Deduce which of the three metals is the strongest reducing agent.

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**(1 mark)**

**(b)** A voltaic cell can be made by joining two half-cells together, such as $Zn/Zn^{2+}$ and $Ni/Ni^{2+}$.

Write a balanced equation for the overall reaction taking place when the two half-cells are connected together, and state which species is undergoing oxidation.

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**(2 marks)**

**(c)** Cell diagrams are a way to represent the redox reactions taking place in voltaic cells.

Write a cell diagram for the overall cell reaction taking place in part (b).

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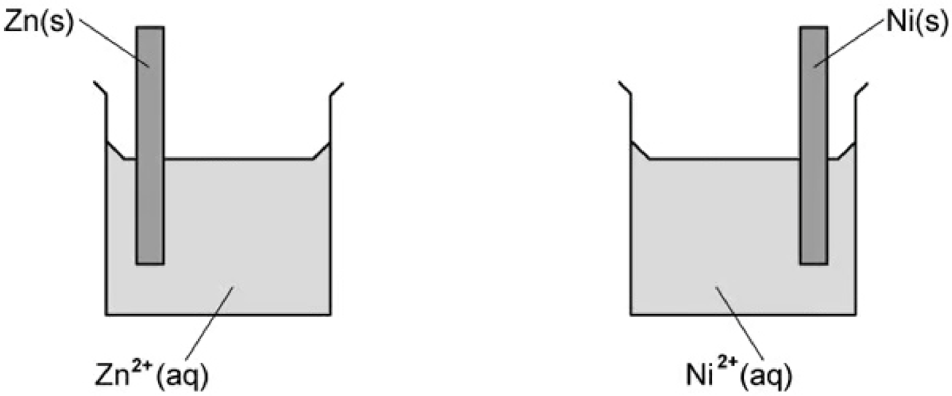
**(1 mark)**

**(d)** Complete the partially labelled diagram in **Figure 1**, of the apparatus used in the voltaic cell in part (b). Show the direction of the movement of the electrons and ions in the cell.

**Figure 1**




**(3 marks)**

# Hard Questions

**1 (a)** A student sets up a titration to determine the amount of iron(II) sulfate in an iron tablet. They titrate the iron(II) sulfate solution with potassium manganate(VII) solution.

i) Write the balanced, ionic half equations to show the reduction of the manganate(VII) ion and the oxidation of the $Fe^{2+}$.

[2]

ii) Use your answers to part (i) to write an overall redox equation for the titration of iron(II) sulfate with potassium manganate(VII) solution.

[1]

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**(3 marks)**

**(b)** The iron(II) sulfate solution is acidified before titration to stop the manganate ion forming unwanted manganese dioxide.

Explain the effect that not acidifying the iron(II) sulfate would have on the final calculation of the estimated mass of iron.

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**(2 marks)**

**(c)** The student dissolved the iron tablet in excess sulfuric acid and made the solution up to 250 $cm^3$ in a volumetric flask. 25.0 $cm^3$ of this solution was titrated with 0.0100 mol $dm^{-3}$ potassium manganate(VII) solution. The average titre was found to be 26.65 $cm^3$ of potassium manganate(VII) solution.

i) Calculate the amount, in moles, of iron(II) ions in the 250 $cm^3$ solution.

[3]

ii) Calculate the mass of iron, in mg, in the tablet.

[2]

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**(5 marks)**

**(d)** Iron sulfate reacts with chromium to produce chromium(III) sulfate, $Cr_2(SO_4)_3$ , and iron

Deduce the overall ionic equation for the reaction occurring

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**(1 mark)**

**2 (a)** Molten potassium bromide can be electrolysed using graphite electrodes.

i) Draw the essential components of this electrolytic cell.

[3]

ii) Identify the products at each electrode.

[2]

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**(5 marks)**

**(b)** State the half equations for the oxidation and reduction processes and deduce the overall cell reaction, including state symbols.

Oxidation half equation ....................

Reduction half equation ..................

Overall equation ...............................

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**(3 marks)**

**(c)** Explain why solid potassium bromide does not conduct electricity.

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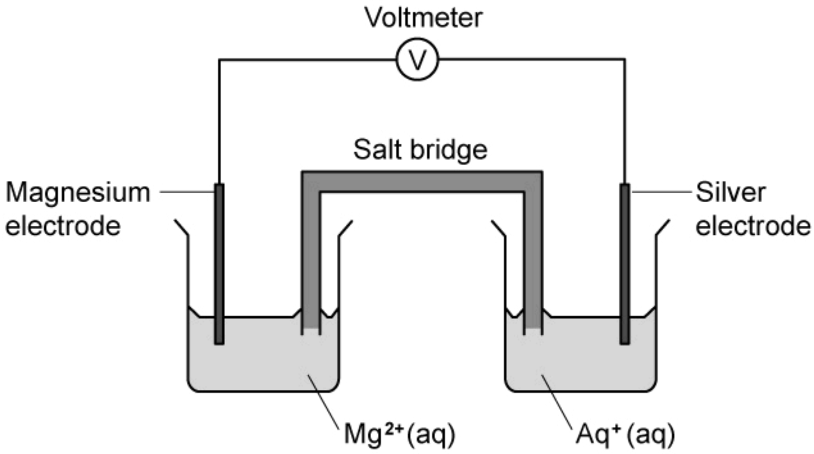
**(1 mark)**

**(d)** A voltaic cell is made from a half-cell containing a magnesium electrode in a solution of magnesium nitrate and a half-cell containing a silver electrode in a solution of silver(I) nitrate.




i) Use section 25 of the data booklet to determine which electrode is positive and to write the equation for the reaction at the positive electrode, including state symbols.

[1]

ii) Compare the processes at the positive electrodes in voltaic and electrolytic cells.

[2]

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**(3 marks)**

**3 (a)** State the oxidation state of phosphorus in the following compounds.

$H_2PO_4^-$ ............................................................

$HPO_3$ ............................................................

$H_3PO_3$ ............................................................

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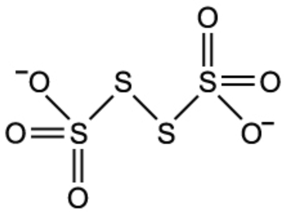
**(3 marks)**

**(b)** The tetrathionate ion is shown below:




i) Determine the oxidation state of sulfur in the ion.

[1]

ii) Justify your answer to part ii).

[1]

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**(2 marks)**

**(c)** Sodium tetrathionate can be formed by reacting sodium thiosulfate, $Na_2S_2O_3$, with iodine.

i) State the balanced symbol equation for this reaction.

[2]

ii) Identify the oxidising agent in this reaction.

[1]

**(3 marks)**

**(d)** Describe the expected observation to show that this reaction had gone to completion.

**(1 mark)**

**4 (a)** A 150.0 $cm^3$ sample of pond water was analysed using the Winkler method to determine its biological oxygen demand (BOD). Initially it took 29.40 $cm^3$ of 0.010 mol $dm^{-3}$ $Na_2S_2O_3$ to react with iodine.

After five days it required 13.70 $cm^3$ of 0.010 mol $dm^{-3}$ $Na_2S_2O_3$ to react with iodine.

The unbalanced equations for the Winker method are shown below.

……$Mn^{2+}$ (aq) + ……$OH^-$ (aq) + $O_2$ (aq) → ……$MnO_2$ (s) + ……$H_2O$ (l)

$MnO_2$ (s) + ……$I^-$ (aq) + ……$H^+$ (aq) → $Mn^{2+}$ (aq) + $I_2$ (aq) + ……$H_2O$ (l)

……$S_2O_3^{2-}$ (aq) + $I_2$ (aq) → $S_4O_6^{2-}$ + ……$I^-$ (aq)

Balance the equations for the Winkler method.

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**(3 marks)**

**(b)** Deduce the reducing agent in the reaction between $S_2O_3^{2-}$ and $I_2$. Justify your answer.

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**(3 marks)**

**(c)** Use the information in part a) and section 6 in your data booklet to determine the initial concentration, in ppm, of oxygen.

(3 marks)

**(d)** Use the information in part a) and section 6 in the data booklet to determine the concentration, in g $dm^{-3}$, of oxygen after five days.

(3 marks)

**(e)** Determine the BOD of the pond water in ppm.

(2 marks)

**5 (a)** 15.00 $cm^3$ of ethanedioic acid, $H_2C_2O_4$ (aq), requires 10.30 $cm^3$ of a 0.250 mol $dm^{-3}$ solution of sodium hydroxide, NaOH (aq), for complete neutralisation using a phenolphthalein indicator for the first permanent colour change.

15.00 $cm^3$ of the same $H_2C_2O_4$ solution required 12.35 $cm^3$ of potassium permanganate solution, $KMnO_4$ (aq), solution for complete oxidation to carbon dioxide and water in the presence of dilute sulfuric acid to further acidify the $H_2C_2O_4$ solution for the first permanent colour change.

i) Using the following equation, calculate the amount, in moles, of $H_2C_2O_4$ (aq).

$$H_2C_2O_4 \text{ (aq)} + 2NaOH \text{ (aq)} \rightarrow Na_2C_2O_4 \text{ (aq)} + 2H_2O \text{ (l)}$$

[2]

ii) Calculate the concentration, in mol $dm^{-3}$, of $H_2C_2O_4$ (aq).

[1]

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**(3 marks)**

**(b)** Deduce the following half equations and overall redox equation for the reaction outlined in part a).

$MnO_4^-$ (aq) to $Mn^{2+}$ (aq) ........................

$H_2C_2O_4$ (aq) to $CO_2$ (g) ..........................

Overall equation ....................................

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**(3 marks)**

**(c)** Calculate the concentration, in mol $dm^{-3}$, of the potassium manganate(VII), $KMnO_4$, solution.

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**(2 marks)**