IB · HL · Chemistry

6 hours 34 questions

Structured Questions

Electron Transfer Reactions

Oxidation & Reduction / Half Equations / Relative Ease of Oxidation & Reduction / Acids with Reactive Metals / Primary Cells / Secondary Cells / Electrolytic Cells / Oxidation of Alcohols / Reduction of Carboxylic Acids, Aldehydes & Ketones / Reduction of Unsaturated Compounds / The Hydrogen Electrode (HL) / Standard Cell Potentials (HL) / Gibbs Energy & Standard Cell Potential (HL) / Electrolysis of...

Easy (10 questions)	/101
Medium (15 questions)	/166
Hard (9 questions)	/72
Total Marks	**/339**

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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 ....

....

....

....

(3 marks)

(b) Chlorine is an *oxidising agent*.

Define *oxidising agent* in terms of electrons.

....

(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^-.

....

....

(2 marks)

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

....

(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.

[1]

ii) Justify your answer to part (i).

[1]

....

....

(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.

[1]

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

ii) Justify your answer.

[2]

....

....

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

....

(1 mark)

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

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

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

[1]

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

[1]

....

....

(2 marks)

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

....

....

(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.

....

(1 mark)

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

....

....

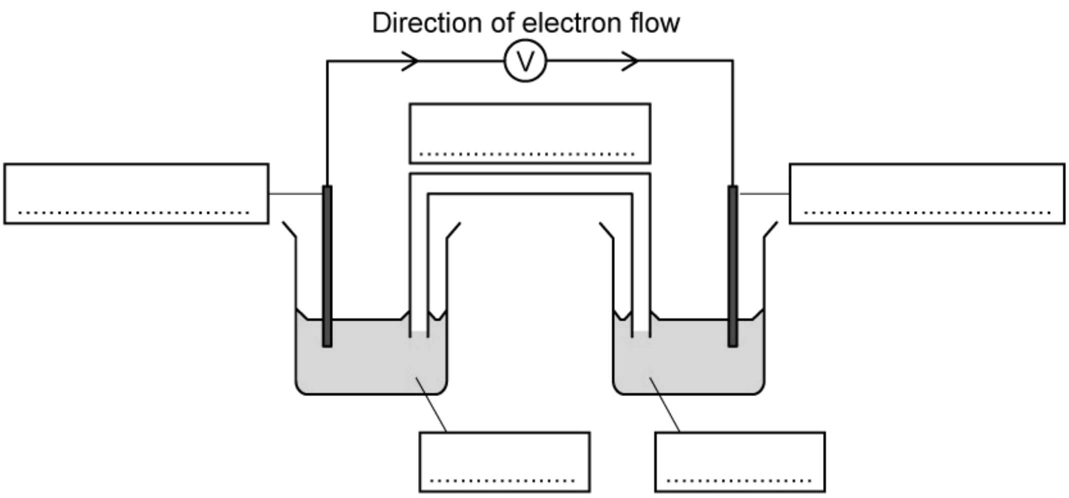
(2 marks)

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

....

(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. 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) State the balanced symbol equations for the complete combustion of propane and propanol.

....

....

....

....

(4 marks)

(b) The following reaction profile produces propanoic acid after three steps:

Propene $\xrightarrow{\text{Step 1}}$ **Propan-1-ol** $\xrightarrow{\text{Step 2}}$ **Propanal** $\xrightarrow{\text{Step 3}}$ **Propanoic acid**

State the reagents and conditions that can be used for steps 2 and 3.

....

....

....

(3 marks)

(c) Using your answer to part (b) to state the colour change for step 2.

....

(1 mark)

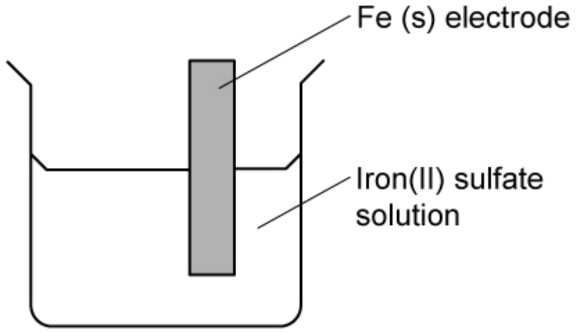
(d) Explain why 2-methylpropan-2-ol will not form a carboxylic acid.

....

....

(2 marks)

6 (a) The image below shows a half cell that can be used to calculate the standard electrode potential of the Fe^{2+} / Fe reaction.



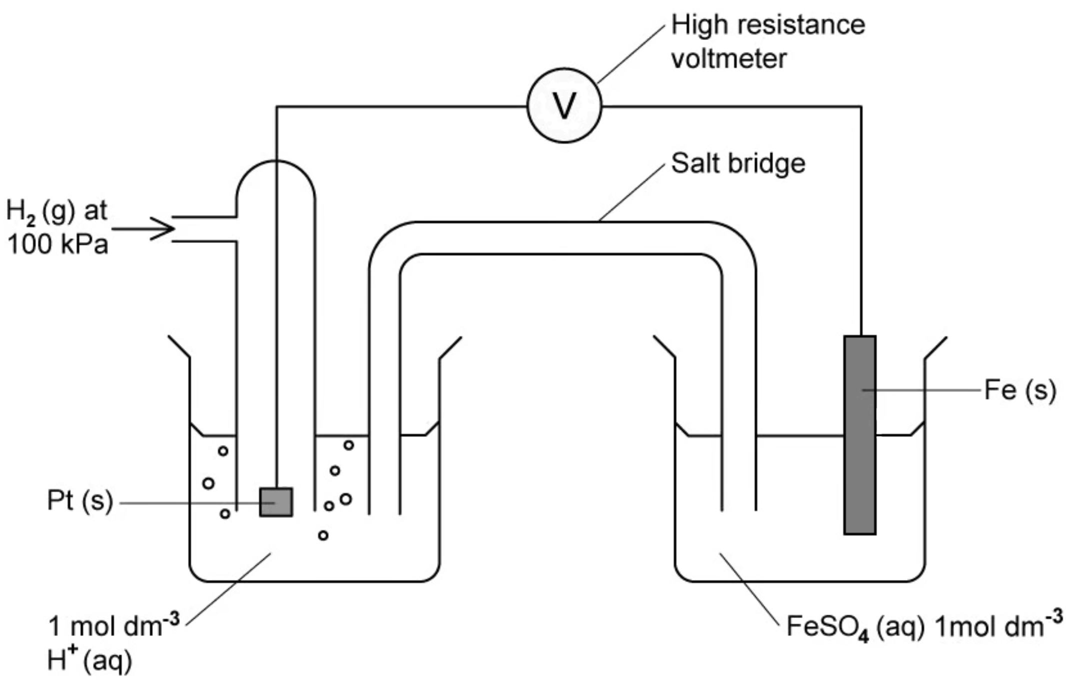
State the half equation, including state symbols, that represents this half cell.

[1]

....

(1 mark)

(b) The electrochemical cell that is used to measure the standard electrode potential of the Fe^{2+} / Fe electrode is shown below.



State the cell representation for the electrochemical cell set up using the standard hydrogen electrode and the Fe^{2+} / Fe electrode.

....

....

....

(3 marks)

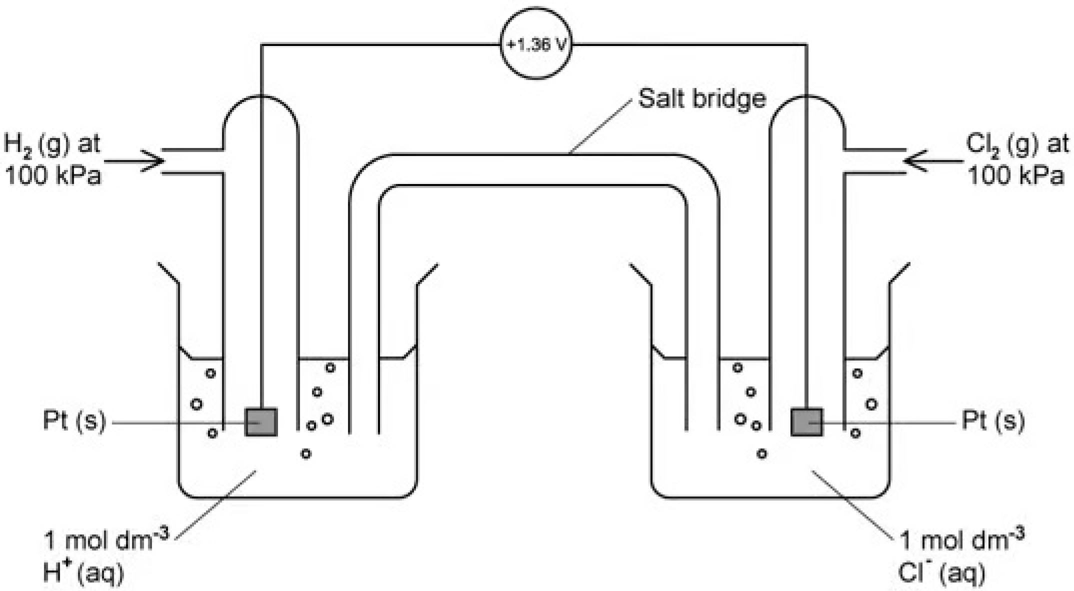
(c) Explain why platinum is used as the electrode for the standard hydrogen electrode.

....

....

(2 marks)

(d) The image shows the electrochemical cell used to measure the standard electrode potential, $E^{\ominus}$, for the Cl_2 / Cl^- half cell.



i) Write the conventional cell representation for this electrochemical cell.

[3]

ii) Determine the standard electrode potential, $E^{\ominus}$, for the Cl_2 / Cl^- half cell.

[1]

(4 marks)

7 (a) State the equation that is required to determine the electromotive force (EMF).

....

(1 mark)

(b) Use section 19 of the data booklet to calculate the electromotive force, in volts, of the following cells.

i) $Zn\ (s)\ |\ Zn^{2+}\ (aq)\ ||\ Cu^{2+}\ (aq)\ |\ Cu\ (s)$

[1]

ii) $Mg\ (s)\ |\ Mg^{2+}\ (aq)\ ||\ Ag^{+}\ (aq)\ |\ Ag\ (s)$

[1]

iii) $Pt\ (s)\ |\ Fe^{2+}\ (aq)\ ,\ Fe^{3+}\ (aq)\ ||\ Cl_2\ (g),\ 2Cl^{-}\ (aq)\ |\ Pt\ (s)$

[1]

....

....

....

(3 marks)

(c) A voltaic cell is constructed using the Ag / Ag^{+} half cell and Pb / Pb^{2+}. Use section 19 of the data booklet to state the following. Include state symbols in your equations.

Half equation for the Ag / Ag^{+} half cell

....

[1]

Half equation for the Pb / Pb^{2+} half cell

....

[1]

Overall equation for the voltaic cell

....

[2]

....

....

....

....

(4 marks)

(d) Use section 19 of the data booklet to determine the electromotive force of the voltaic cell outlined in part c).

....

(1 mark)

8 (a) Use section 19 of the data booklet to answer the following questions about the electrolysis of **dilute** sodium chloride solution using inert electrodes.

i) State the equations to generate the ions present in solution.

[2]

ii) Predict the product at the anode.

[2]

iii) Predict the product at the cathode.

[2]

(6 marks)

(b) Use section 19 of the data booklet to predict the products at the anode and cathode for the electrolysis of copper sulfate with inert electrodes.

i) State the equations to generate the ions present in solution.

[2]

ii) Predict the product at the anode.

[2]

iii) Predict the product at the cathode.

[2]

(6 marks)

(c) Use your answer to part b) to write an overall equation for the electrolysis of copper sulfate using inert electrodes.

(3 marks)

(d) The inert electrodes for the electrolysis of copper sulfate are replaced by copper electrodes. State the half equations that occur at the anode and cathode that occur with copper electrodes.

(2 marks)

9 (a) State the value above which the value for the standard electrode potential, E^{θ}_{cell} value, indicates a reaction is spontaneous.

....

(1 mark)

(b) The spontaneous reaction between zinc and copper in a voltaic cell is shown below

$$Ni\ (s) + Cu^{2+}\ (aq) \rightarrow Ni^{2+}\ (aq) + Cu\ (s) \quad E^{\theta}_{cell} = +0.60\ V$$

Use sections 1 and 2 of the data booklet to determine the free energy change, ΔG^{θ}, for the reaction in kJ mol^{-1}.

....

....

(2 marks)

(c) Use section 19 of the data booklet to determine if the reaction shown is spontaneous at standard conditions

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

....

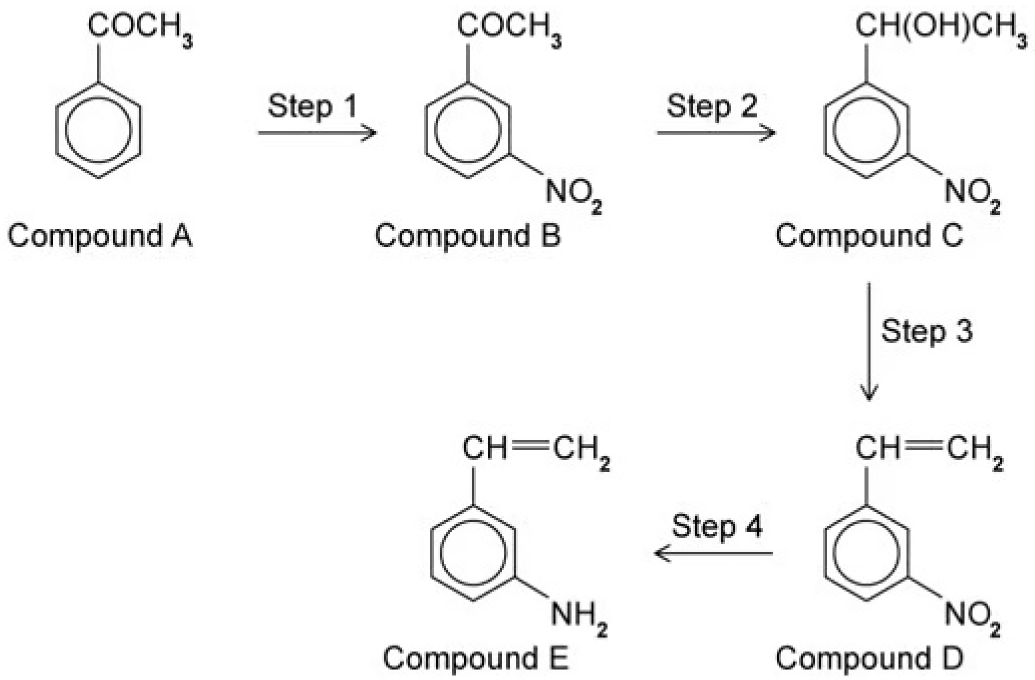
....

....

....

(4 marks)

10 (a) The synthesis of 3-aminstyrene is shown below:



i) Give the reagent needed in Step 1.

[1]

ii) State the name of the functional groups in Compound B.

[2]

....

....

....

(3 marks)

(b) This question is about Step 2.

i) Give the reagent needed.

[1]

ii) Name the type of reaction taking place.

[1]

....

(2 marks)

(c) Step 3 is a dehydration reaction. Outline a chemical test that could distinguish between Compound C and the product of Step 3, Compound D.

(2 marks)

(d) This question is about Step 4.

i) State the name of the reagent(s) and conditions needed in Step 4.

[2]

ii) Identify the type of reaction taking place.

[1]

(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.

[2]

ii) Deduce the oxidation number of chlorine in all of the chlorine-containing reactants and products

[1]

....

....

....

(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?

[1]

ii) What is the oxidizing agent in this reaction? Explain your answer.

[2]

iii) Why is the hazard of the toxic chlorine gas being produced greater than the hazard of hydrazine?

[1]

....

....

(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.

[1]

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

[1]

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

[1]

iv) Deduce the overall redox equation for this reaction.

[2]

(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.

....

....

(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.

....

....

(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.

[1]

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

[1]

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

[2]

....

....

....

....

(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.

....

....

....

(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.

....

....

....

....

(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.

[2]

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

[2]

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....

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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.

[1]

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

[2]

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

[2]

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

[2]

(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.

[1]

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

[1]

....

....

(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*.

....

(1 mark)

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

....

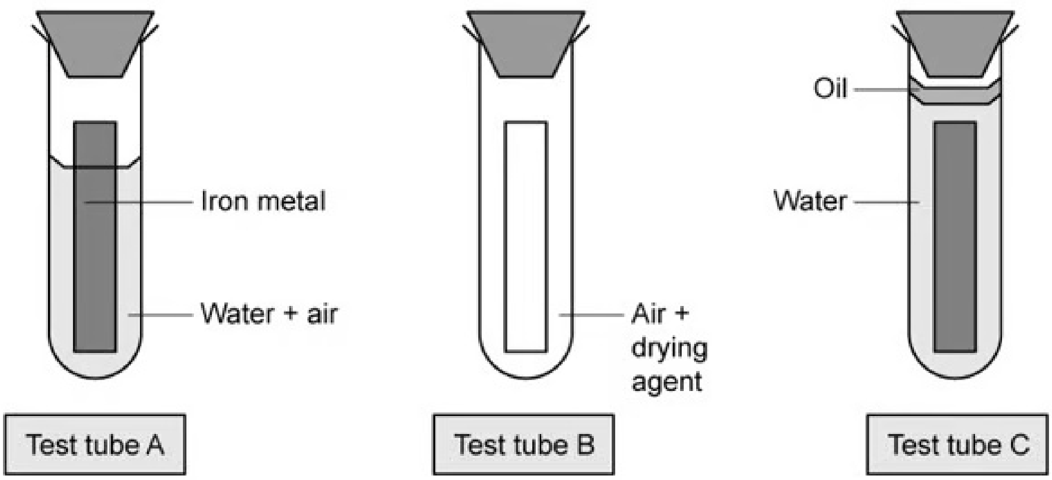
....

(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.

....

(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 ClF	

(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.

[1]

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

[2]

....

....

....

(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.

[2]

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

[2]

....

....

....

....

(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.

....

....

....

(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.

....

(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.

....

....

(2 marks)

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

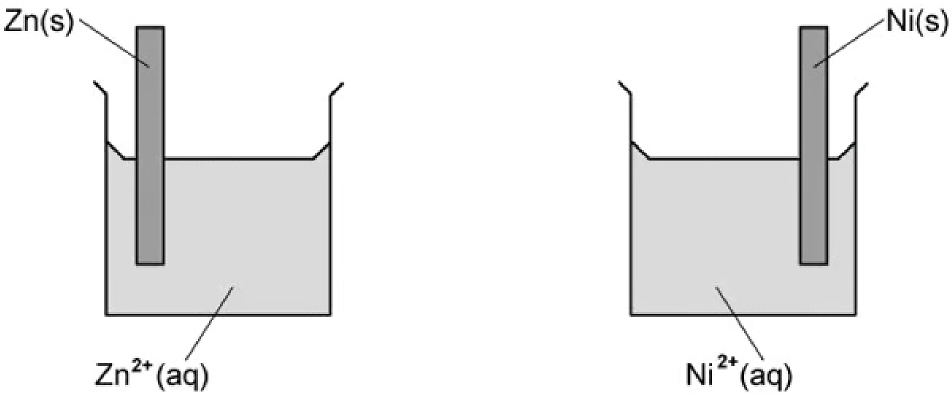
Write a cell diagram for the overall cell reaction taking place in part (b).

....

(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)

7 (a) Ethene, C_2H_4, can be made into a number of useful compounds. A reaction sequence for this is shown below:

$$C_2H_4 \xrightarrow{\text{Step 1}} C_2H_5Cl \xrightarrow{\text{Step 2}} C_2H_6O \xrightarrow{\text{Step 3}} C_2H_4O$$

i) Name the type of reaction shown in step 1.

[1]

ii) Write an equation, using structural formulas, for the reaction in step 2 in which C_2H_5Cl reacts with aqueous NaOH to form C_2H_6.

[1]

....

....

(2 marks)

(b) The product of step 2 can undergo combustion.

i) Write a balanced equation for the *complete* combustion of the product of step 2.

[1]

ii) Write a balanced equation for the *incomplete* combustion of the product of step 2.

[1]

....

....

(2 marks)

(c) Give the reagents and conditions needed to carry out step 3.

....

....

(2 marks)

(d) The product of step 2 has a higher boiling point than the product of step 3.

State the names of the products of step 2 and 3, and explain the difference in their boiling points.

(3 marks)

8 (a) Two isomeric compounds are shown below in **Figure 1**.

Figure 1

A

B

i) State the name of each isomer.

[1]

ii) Suggest a chemical reagent to distinguish between these isomers and deduce the type of reaction taking place.

[2]

iii) State the observations made in each case.

[2]

.....

.....

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.....

(5 marks)

(b) Compound B, $CH_3CH(CH_3)CH(OH)COOH$, can be oxidized into compound C.

i) Deduce the half-equation for the conversion of compound B into C.

[1]

ii) The half equation for the oxidation reaction using acidified potassium dichromate(VI) is as follows:

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \quad 2Cr^{3+}(aq) + 7H_2O\ (l)$$

Deduce the overall redox equation for the conversion of B into C.

[2]

....

....

....

(3 marks)

(c) The same reaction in part (b) can be used to oxidize ethanol into ethanal or ethanoic acid, depending on the reaction conditions.

Outline how the reaction conditions can be changed to produce ethanal or ethanoic acid.

....

....

(2 marks)

9 (a) Some standard electrode potential data are shown in **Table 1** which you will need to answer the following questions.

Table 1

Half-equation	***E*$^{\ominus}$ / V**
Cu^{2+} (aq) + $2e^-$ ⇌ Cu (s)	+0.34
Ni^{2+} (aq) + $2e^-$ ⇌ Ni (s)	-0.25
Fe^{3+} (aq) + e^- ⇌ Fe^{2+} (aq)	+0.77
Sn^{2+} (aq) + $2e^-$ ⇌ Sn (s)	−0.14
Fe^{2+} (aq) + $2e^-$ ⇌ Fe (s)	−0.44

Deduce the species from **Table 1** that is the weakest oxidising agent. Explain your choice.

....

....

(2 marks)

(b) Give the conventional representation of the cell that is used to measure the standard electrode potential of copper/copper(II) ions as shown in **Table 1** in part (a).

....

....

(2 marks)

(c) A voltaic cell is made from nickel in a solution of nickel(II) chloride and copper in a solution of copper(II) sulfate.

Calculate the EMF of this cell using the values given in **Table 1** in part (a).

....

(1 mark)

(d) Two half-cells, involving species in **Table 1**, are connected together to give a cell with an EMF = +0.30 V

i) Determine which two half equations produce this EMF using the data from **Table 1** and write the overall equation for the reaction.

[2]

ii) Suggest the half-equation for the reaction that occurs at the positive electrode (cathode).

[1]

....

....

....

(3 marks)

10 (a) Aqueous copper(II) sulfate can be electrolysed using passive or active electrodes. Passive electrodes can be made of platinum and active electrodes from copper.

Draw a labelled diagram of an electrolytic cell for this process using platinum electrodes and identify in which direction electrons flow.

....

....

(2 marks)

(b) Write the half equations taking place at each electrode in part a), including state symbols, and state what is seen at each electrode.

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(4 marks)

(c) Write the half equations taking place at each electrode when using copper electrodes, including state symbols, and state what is seen at each electrode.

....

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(4 marks)

(d) State what happens to the colour and acidity of the electrolyte when using platinum and copper electrodes in the electrolysis of aqueous copper(II) sulfate.

....

....

(4 marks)

11 (a) State the conditions under which the EMF of a redox reaction will be spontaneous.

....

(1 mark)

(b) Using Sections 1 & 19 of the Data Booklet, calculate Δ*G* for the following reaction and state whether the reaction is spontaneous under standard conditions.

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

....

....

....

(3 marks)

(c) Suggest, with a reason, how a non-spontaneous reaction could be made spontaneous.

....

....

(2 marks)

(d) Using **Table 2**, predict and write overall equations for all the spontaneous reactions.

Table 2

Half-equation	**E^{θ} / V**
$Ag^{+}(aq) + e^{-} \rightleftharpoons Ag(s)$	+0.80
$\frac{1}{2} I_2(aq) + e^{-} \rightleftharpoons I^{-}(aq)$	+0.54
$Sn^{2+}(aq) + 2e^{-} \rightleftharpoons Sn(s)$	−0.14

....

....

(3 marks)

12 (a) Metals coatings on other metals can be achieved using electroplating. Three beakers containing solutions of $Sn(NO_3)_4$, $Co_2(SO_4)_3$, $Pb(NO_3)_2$, were set up as electrolytic cells and used to electroplate the metals. The same amount of current was passed through the cells for the same length of time.

State and explain in which cell would the greatest amount of metal be produced and identify the electrode where the metals are deposited.

....

....

....

....

(4 marks)

(b) Apart from current and time, identify two factors that influence the amount of cobalt deposited in the $Co_2(SO_4)_3$ cell.

....

....

(2 marks)

(c) State **two** reasons why electroplating of metals is carried out.

....

....

(2 marks)

(d) A nickel teaspoon is electroplated with silver using sodium argentocyanide. Predict the mass changes at each electrode.

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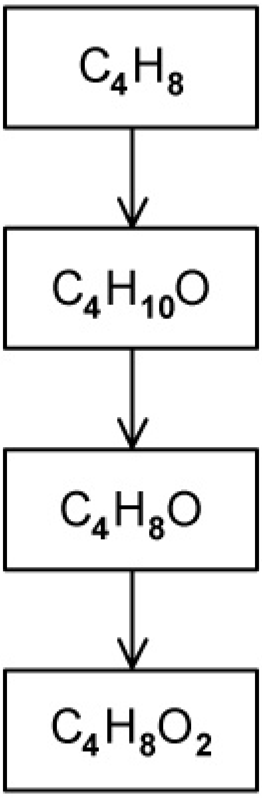
(1 mark)

13 (a) An organic reaction sequence is shown below.



State the IUPAC names of the four substances in the sequence.

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(4 marks)

(b) Classify the reactions in (a) and give the names of the reagents in each step.

....

....

....

(6 marks)

(c) Give the reaction conditions for step 3 in (a)

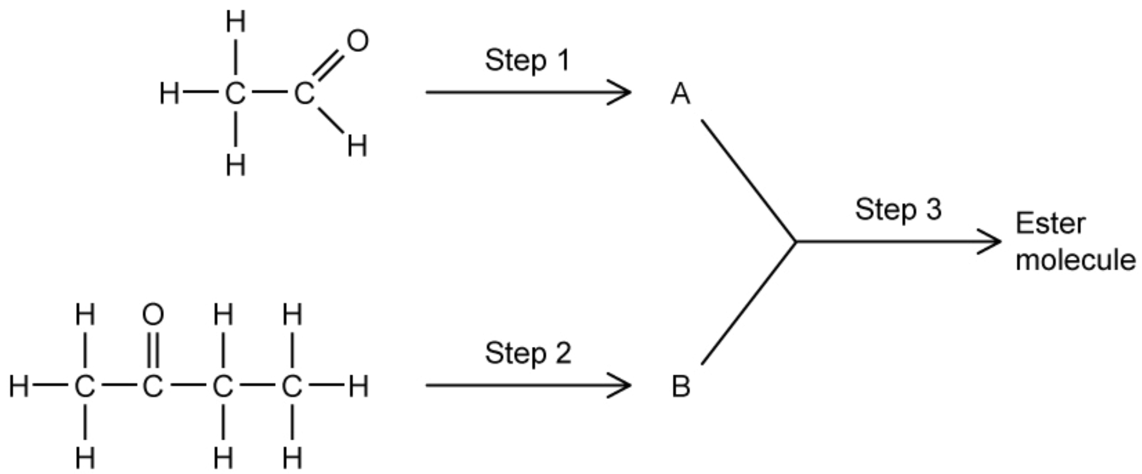
(1 mark)

(d) Draw a displayed formula of an isomer of $C_4H_{10}O$ that gives two signals in an ^1H NMR spectrum.

(1 mark)

14 (a) The following reaction pathway is used to produce Compounds **A** and **B**, which when reacted together, form a branched ester molecule, Compound **C**.

Suggest suitable reagents and conditions for the synthesis of Compound **A** via Step 1 and give the name for this type of reaction.



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(3 marks)

(b) In order for the ester to be produced, the ketone in part (a) must be converted to another compound, **B**.

i) Name and draw the structure of the molecule that is produced from Step 2.

ii) Give the name of the type of reaction that is involved in Step 2 and suggest suitable reagents and conditions for the process.

....

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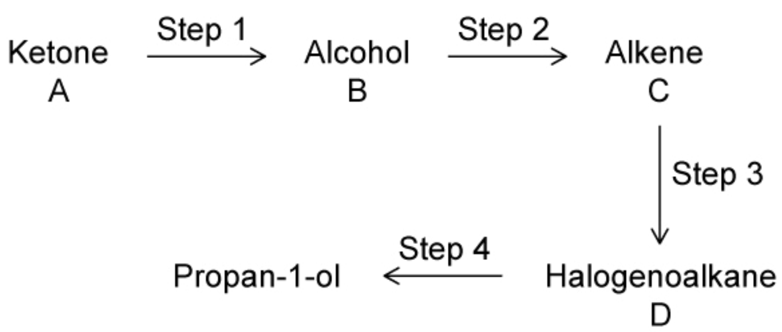
....

(4 marks)

(c) Outline how ethanol can be synthesised from ethane in two steps. State the reaction conditions and reagents and name the type of reaction taking place.

(6 marks)

(d) The four step synthesis to form propan-1-ol from a ketone is outlined below.



i) Give the names of four possible substances **A** to **D**

ii) Give the reagents and conditions for Step 4.

(3 marks)

15 (a) Propanal is a versatile organic building block used in the synthesis of plastics and rubber chemicals.

Propanal can be produced from propanoic acid in the following two-step reaction.

Propanoic acid $\xrightarrow{\textbf{Step 1}}$ **Propan-1-ol** $\xrightarrow{\textbf{Step 2}}$ **Propanal**

State the reaction type, including suitable reagents, for Steps 1 and 2.

(2 marks)

(b) Suggest why it is not possible to convert propanoic acid directly to propanal using the reagent you identified for Step 1 in (a).

(1 mark)

(c) Explain why Step 2, in (a), is completed by distillation.

(1 mark)

(d) Identify, explain your reasoning, which of the three organic compounds, from the reaction scheme in (a), would be distilled first.

(2 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]

(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.

(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

....

(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]

....

....

....

....

....

(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 ....

....

....

....

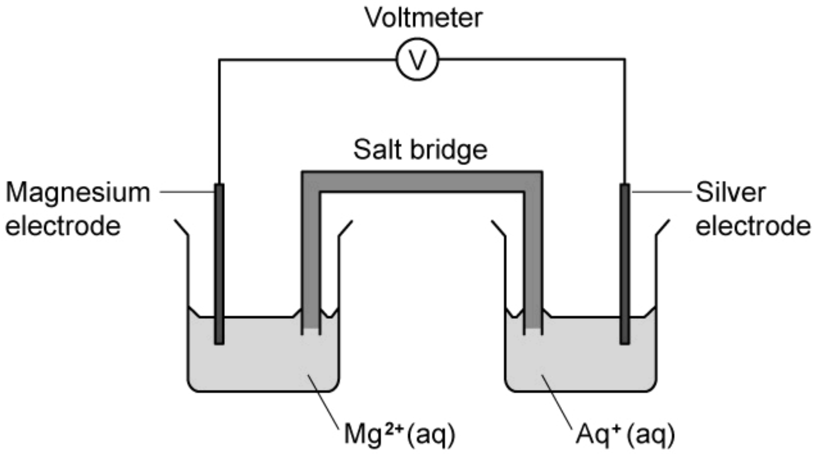
(3 marks)

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

....

(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 ....

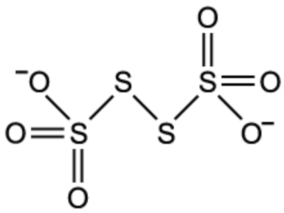
....

....

....

(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]

....

....

(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) 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]

....

....

....

(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 ....

....

....

....

(3 marks)

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

....

....

(2 marks)

5 (a) Use section 19 of the data booklet to draw the electrochemical cell for the feasible reaction of Ag / Ag^{+} and Al / Al^{3+}. Write the conventional representation, including state symbols, for this cell.

(3 marks)

(b) Write the conventional representation, including state symbols, for this cell.

(1 mark)

(c) Explain why the salt bridge connecting the silver and aluminum electrodes cannot be made with potassium chloride solution.

(2 marks)

(d) The silver half cell is replaced with a magnesium half cell. Deduce the reading on the voltmeter.

(2 marks)

6 (a) Use section 19 of the data booklet and the information below to determine if the following reaction is feasible at 298 K.

$$2KMnO_4 \text{ (aq)} + 5H_2O_2 \text{ (aq)} + 6HCl \text{ (aq)} \rightarrow 2MnCl_2 \text{ (aq)} + 8H_2O \text{ (l)} + 5O_2 \text{ (g)} + 2KCl \text{ (aq)}$$

$$O_2 \text{ (g)} + 2H^+ + 2e^- \rightleftharpoons H_2O_2 \quad E^\theta = 0.68 \text{ V}$$

....

....

(2 marks)

(b) The reaction of copper oxide and sulfuric acid is shown below. Use section 19 of the data booklet to explain why the reaction is thermodynamically feasible.

$$CuO \text{ (aq)} + H_2SO_4 \text{ (aq)} \rightarrow CuSO_4 \text{ (aq)} + H_2O \text{ (l)}$$

....

....

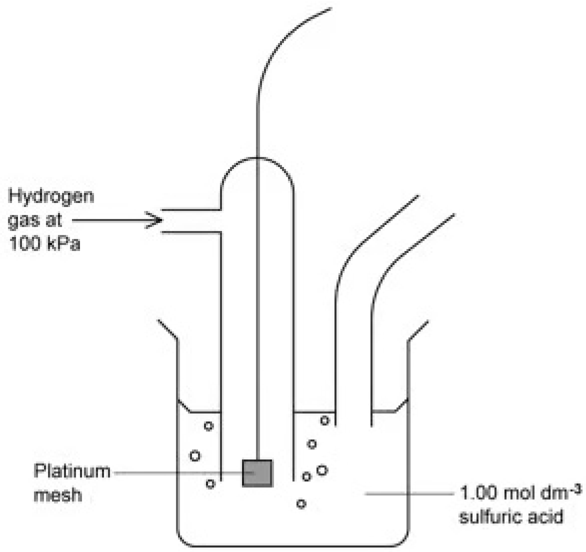
(2 marks)

(c) Suggest a reason why the reaction does not occur despite being thermodynamically feasible.

....

(1 mark)

7 (a) Explain why the following does not represent the standard hydrogen electrode.

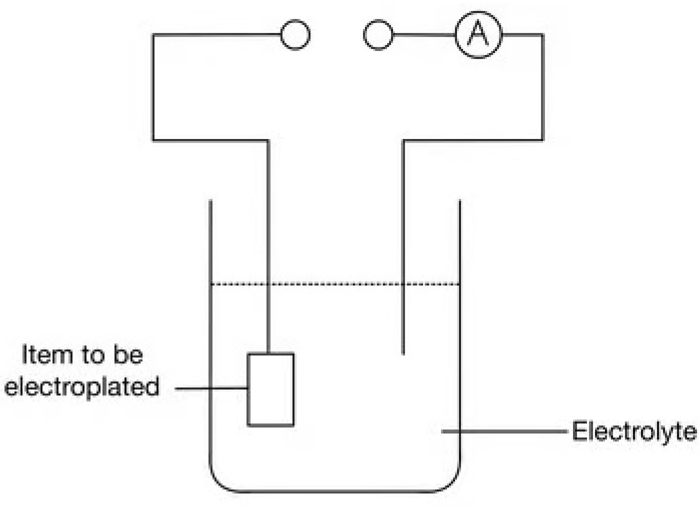


(2 marks)

(b) The standard electrode potential for Zn^{2+} (aq) + $2e^-$ → Zn (s) is –0.76 V. State the meaning of the minus sign in the value of –0.76 V.

(1 mark)

(c) Zinc coating on metals serves as physical protection which prevents rust from affecting the underlying metal surface. This is achieved by electroplating.



i) Suggest a suitable solution to act as the electrolyte during zinc electroplating.

[1]

ii) Complete the diagram by labelling the polarity of the power source by using a + and - sign.

[1]

....

....

(2 marks)

8 (a) Using section 19 of the data booklet deduce the full equation for the $Cr_2O_7^{2-}$ (aq) / Cr^{3+} (aq) and Br_2 (l) / Br^- (aq) cell.

....

(1 mark)

(b) Determine the value for $E^\ominus$cell value for the cell outlined in part a).

....

(1 mark)

(c) Use your answer to part b) and sections 1 and 2 of the data booklet to determine whether the reaction in part a) reaction is spontaneous.

....

(1 mark)

(d) An electrochemical cell has a free energy change of -14.475 kJ mol^{-1}. Use the information in the table to determine the cell representation of the electrochemical cell.

Electrode half-equation	$E^\ominus$ **/ V**
Ag^+ (aq) + e^- ⇌ Ag (s)	+0.80
Li^+ (aq) + e^- ⇌ Li (s)	-3.04
ClO_2 (aq) + e^- ⇌ ClO_2^- (aq)	+0.95
H_2O (l) + e^- ⇌ ½H_2 (g) + OH^- (aq)	-0.83
Fe^{3+} (aq) + e^- ⇌ Fe^{2+} (aq)	+0.77

....

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(2 marks)

9 (a) Aqueous sodium tetrahydridoborate, $NaBH_4$, is a common reducing agent.

State the IUPAC name of the two isomers with the formula C_3H_6O that can be reduced by aqueous $NaBH_4$.

....

....

(2 marks)

(b) State the IUPAC name of **two** non-cyclic isomers with the formula C_3H_6O that cannot be reduced by aqueous $NaBH_4$.

....

....

(2 marks)

(c) When $NaBH_4$ is used as a reducing agent followed by the addition of acid, the reduction products of ketones can exhibit optical isomerism, while the reduction products of aldehydes cannot.

i) Classify the reduction products of aldehydes and ketones.

[2]

ii) Explain why the reduction products of ketones can exhibit optical isomerism, while the reduction products of aldehydes cannot.

[2]

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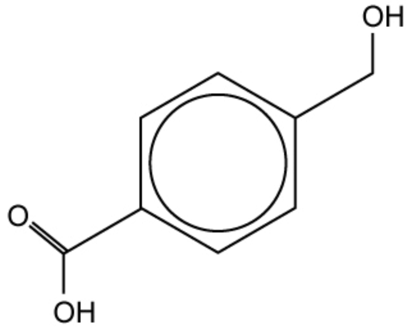
....

....

(4 marks)

(d) Deduce the structure when the following compound is reduced using $NaBH_4$.



(1 mark)