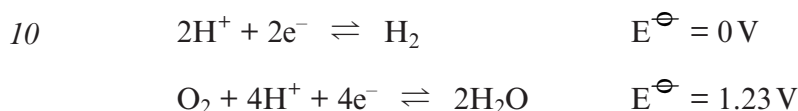**3.** *Read the passage below and then answer questions (a) to (d) in the spaces provided.*

Hydrogen Fuel Cells

1 Although fuel cells have been around since 1839, it took another 120 years until NASA demonstrated some of their potential applications when providing power during space flights.

5 A fuel cell works like an electrochemical cell (battery) but does not run down or need recharging. It will produce electricity and heat as long as fuel (hydrogen) is supplied. A fuel cell consists of two electrodes—an anode where oxidation occurs and a cathode for reduction—sandwiched around an electrolyte. Replacing the salt bridge of conventional electrochemical cells, several electrolyte systems have been tried such as phosphoric acid or a solid electrolyte based on polymeric fluorocarbons. The relevant electrode potentials are

10 $$2H^+ + 2e^- \; \rightleftharpoons \; H_2 \qquad\qquad E^{\ominus} = 0\,V$$

$$O_2 + 4H^+ + 4e^- \; \rightleftharpoons \; 2H_2O \qquad\qquad E^{\ominus} = 1.23\,V$$

Hydrogen is fed to the anode, and oxygen (air) to the cathode. Activated by a catalyst, usually involving a layer of platinum and carbon a few nanometres thick, hydrogen atoms separate into protons and electrons, which take different paths to the cathode. The electrons
15 go through an external circuit, creating a flow of electricity. The protons migrate through the electrolyte. Fuel cells can be used to power vehicles or to provide electricity and heat to buildings.

A significant barrier to using fuel cells in vehicles is hydrogen storage. Most fuel-cell vehicles powered by hydrogen store the hydrogen as a compressed gas in pressurized tanks. Due to
20 the low energy density of hydrogen, it is difficult to store enough hydrogen onboard to allow vehicles to travel the same distance as petrol-powered vehicles.

A potentially energy-dense water-based fuel is based on sodium tetrahydridoborate(III) (30% by mass $NaBH_4$ in water). A catalyst induces rapid hydrogen production

$$NaBH_4 \; + \; 2H_2O \; \rightarrow \; 4H_2 \; + \; NaBO_2 \qquad\qquad \Delta H^{\ominus} = -300\,kJ\,mol^{-1}$$

25 and pure humidified H_2 is delivered to the engine or fuel cell. The exothermic reaction requires no heat input and sodium borate, $NaBO_2$, can be recycled into $NaBH_4$.

– End of passage –

(a)   State the function performed by both the salt bridge in an electrochemical cell and the electrolyte in a fuel cell. *(lines 6-7)* [2]

....

....

(b)   (i)   Explain why the $2H^+ + 2e^- \rightleftharpoons H_2$ electrode has an electrode potential of zero. *(line 10)* [1]

....

....

(ii)   Calculate the EMF of the hydrogen fuel cell. *(lines 10-11)* [1]

....

....

(iii)   Give **one** reason why the EMF calculated in *(b)*(ii) is not attained in practice, with 0.7 V being a typical value for a fuel cell. [1]

....

....

(iv)   Write a balanced equation for the overall reaction which occurs in the cell. *(lines 10-11)* [1]

....

....

(v)   Given that $\Delta H_f^{\ominus} H_2O(l) = -285.8 \, kJ \, mol^{-1}$, calculate the enthalpy change, $\Delta H^{\ominus}$, for the equation in *(b)*(iv). [1]

....

....

(c) (i) State **one** disadvantage, mentioned in the passage, of using hydrogen fuel cells to power vehicles. *(lines 18-21)* [1]

....

....

(ii) Give a second disadvantage, not mentioned in the passage, of using hydrogen as a fuel in vehicles. [1]

....

....

(iii) State **one** advantage of using a hydrogen fuel cell compared to the combustion of petrol. [1]

....

....

(d) When 1 kg of the water-based fuel (30% $NaBH_4$ by mass) is reacted to produce hydrogen, calculate *(lines 22-26)*

(i) the mass, and hence the number of moles, of $NaBH_4$ in 1 kg of the water-based fuel, [2]

....

....

(ii) the energy given out (kJ) by 1 kg of the water-based fuel, [1]

....

....

(iii) the volume of hydrogen gas produced. [2]

[Assume 1 mol H_2 gas occupies a volume of 24 dm^3]

....

....

....

....

Total [15]

Turn over.

(b) Both boron nitride, BN, and carbon, C, form hexagonal graphite-type structures. Explain why
   • BN and C can both adopt the same hexagonal structure;
   • both BN and C exhibit lubricating properties;
   • C is an electrical conductor but BN is an insulator at room temperature.   [6]

(QWC) [2]

Total [20]

5. (a) *Bordeaux Mixture* is one of the earliest fungicides, first used about 1885. It can be prepared by mixing copper sulfate solution with excess limewater (calcium hydroxide solution).

   (i) State what you would observe when copper sulfate solution is mixed with limewater.   [2]

   (ii) Write an equation for the reaction that occurs.   [1]

   (b) A sample of *Bordeaux Mixture* was analysed to determine its copper content. Firstly, it was reacted with excess potassium iodide

$$2Cu^{2+} + 4I^- \rightarrow 2CuI + I_2$$
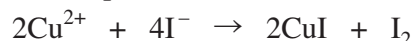
and the iodine produced was then titrated against sodium thiosulfate solution.

$$I_2 + 2Na_2S_2O_3 \rightarrow 2NaI + Na_2S_4O_6$$
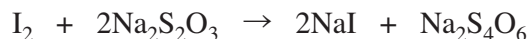
   (i) Name the indicator used for the titration and state the colour change at the end-point.   [2]

   (ii) If a 31.2 g sample of *Bordeaux Mixture* required 12.25 cm^3 of sodium thiosulfate solution with concentration 0.100 mol dm^{-3} Na$_2$S$_2$O$_3$ to react with the liberated iodine, calculate the mass of copper in the sample and hence the % Cu by mass in *Bordeaux Mixture*. Your answers should be given to **three** significant figures.   [3]

   (c) Copper can exist as Cu^{2+} or Cu$^+$ compounds.

   (i) Write the full electron configurations for Cu^{2+} ions **and** Cu$^+$ ions.   [2]

   (ii) Explain why most Cu^{2+} compounds are coloured blue in the presence of water.   [4]

   (iii) Briefly explain why most Cu$^+$ compounds are colourless or white.   [1]

   (d) (i) State what would be observed, and give equations for any reactions, when tetrachloromethane, CCl$_4$, and silicon(IV) chloride, SiCl$_4$, are separately added to water.   [3]

   (ii) Explain why lead forms a solid chloride PbCl$_2$, but the corresponding CCl$_2$ and SiCl$_2$ are too unstable to exist.   [2]

Total [20]

SECTION A

*Answer **all** questions in the spaces provided.*

1. Chromium and aluminium both form amphoteric compounds.

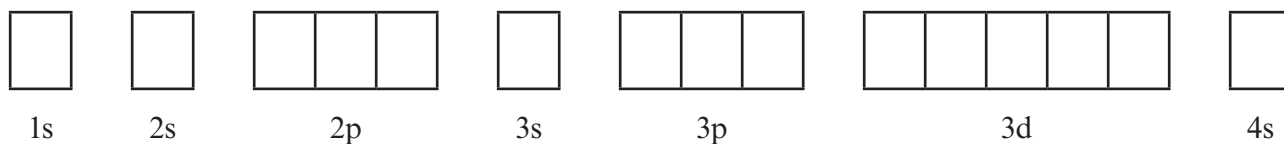   *(a)* State what is meant by the term *amphoteric*.   [1]
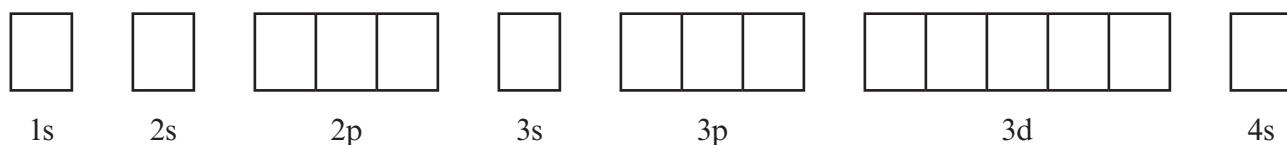
    ....

    ....

    (b) Use arrows in boxes to show the electronic structures of the chromium atom and the Cr^{3+} ion.   [2]

    Chromium atom, Cr

☐	☐	☐☐☐	☐	☐☐☐	☐☐☐☐☐	☐
1s	2s	2p	3s	3p	3d	4s

    Chromium(III) ion, Cr^{3+}

☐	☐	☐☐☐	☐	☐☐☐	☐☐☐☐☐	☐
1s	2s	2p	3s	3p	3d	4s

    (c) When sodium hydroxide solution is added to a solution of potassium dichromate(VI), $K_2Cr_2O_7$, the following reaction occurs.

    $$Cr_2O_7^{2-} \ + \ 2OH^- \ \longrightarrow \ 2CrO_4^{2-} \ + \ H_2O$$
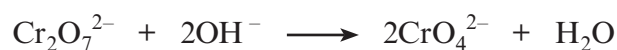
    (i) State the colour change that is seen.   [1]

    ....

    (ii) Use the oxidation states of chromium to show that this is not a redox reaction. [2]

    ....

    ....

2. Fuel cells have been proposed as an alternative method of providing energy for vehicles. These use chemical reactions within electrochemical systems to generate electricity.

(a) A typical fuel cell uses hydrogen as a fuel and reacts this with oxygen. The two half-equations for the processes occurring at the electrodes are given in the table below.

Half-equation	$E^{\ominus}$ / V
$2H^+ + 2e^- \rightleftharpoons H_2$	0.00
$\frac{1}{2}O_2 + 2H^+ + 2e^- \rightleftharpoons H_2O$	1.23

(i) Write an equation for the overall reaction occurring. [1]

....

(ii) Give **one** benefit of the use of fuel cells as a replacement for traditional vehicle energy sources. [1]

....

....

(iii) The same reaction as above can be undertaken in a traditional electrochemical cell, such as the one below. Name the parts labelled **A-C**. [3]

B ....

hydrogen, 1 atm →

A ....

oxygen, 1 atm

1 mol dm^{-3} H$^+$(aq) →

← 1 mol dm^{-3} H$^+$(aq)

C ....
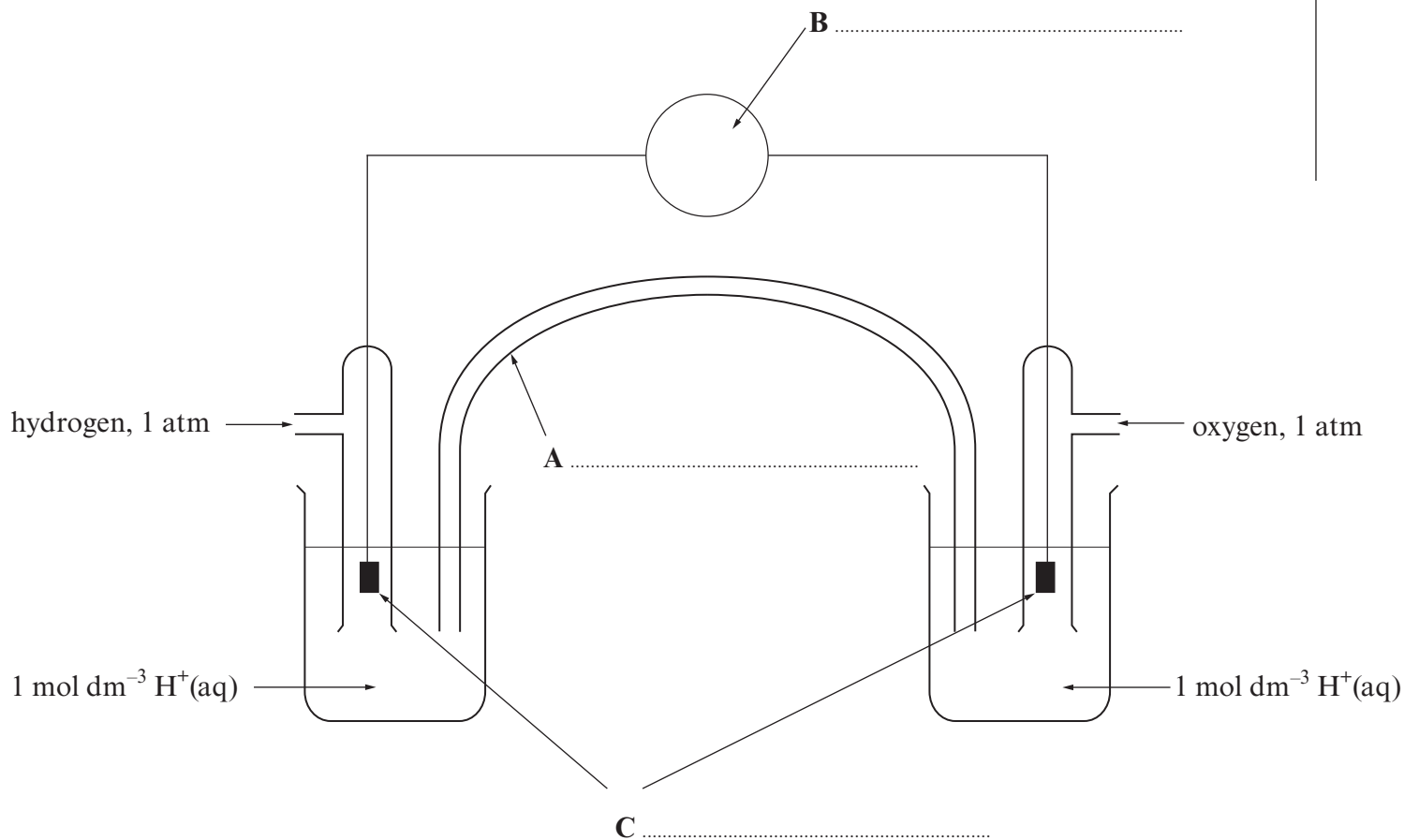
SECTION B

*Answer **both** questions in the separate answer book provided.*

4. *(a)* (i) State what is meant by the term *transition element*. [1]

(ii) Explain why both iron and copper are classed as transition elements, whilst zinc is not. [1]

(b) Transition elements such as copper frequently form coloured complexes. Copper(II) complexes are usually blue, but the exact colour can vary, with $[Cu(H_2O)_6]^{2+}$ being pale blue and $[Cu(NH_3)_4(H_2O)_2]^{2+}$ being royal blue. Copper(I) complexes are usually colourless.

Explain why transition metal complexes are usually coloured. Your answer should include details of:

• The origin of colour in transition metal complexes;

• Why the copper(II) species above are coloured blue;

• Why the colours seen in different copper(II) complexes are different;

• Why copper(I) complexes do not form coloured compounds. [6]

(QWC) [2]

(c) Iron is usually extracted from iron(III) oxide, Fe_2O_3, in a blast furnace using carbon monoxide, CO, as a reducing agent, releasing metallic iron and the gas carbon dioxide.

(i) Write the overall equation for this reaction. [1]

(ii) Explain in terms of oxidation states why carbon monoxide is considered to be the reducing agent in this reaction. [2]

(iii) Explain why carbon monoxide, CO, can be used as a reducing agent but the corresponding oxide of lead, PbO, cannot. [2]

(d) The iron content of an alloy can be determined by a redox titration using acidified potassium dichromate(VI) solution, $K_2Cr_2O_7$. A piece of alloy of mass 1.870 g was dissolved completely in acid to form Fe^{2+} ions, and the solution made up to 250.0 cm^3. A 25.00 cm^3 sample of this solution was titrated against acidified $K_2Cr_2O_7$. This required 23.80 cm^3 of $K_2Cr_2O_7$ solution of concentration 0.0200 mol dm^{-3} for complete reaction.

   (i) The half-equations for the processes occurring are:

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

$$Fe^{3+} + e^- \longrightarrow Fe^{2+}$$
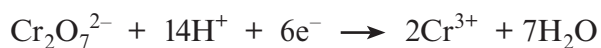
   Write an **ionic** equation for the reaction between Fe^{2+} ions and $Cr_2O_7{}^{2-}$ ions in acid solution. [1]

   (ii) Calculate the number of moles of Fe^{2+} ions present in the 25.00 cm^3 sample used in the titration. [2]

   (iii) Calculate the percentage of iron in the original alloy sample. [2]

Total [20]

5. *(a)* Give a current use for a named compound of chlorine. [1]

(b) Chlorine gas, Cl_2, is used in the industrial preparation of bromine, Br_2. Sea water contains small amounts of bromide ions and by bubbling chlorine gas through the sea water these can be converted to Br_2.

(i) Write an ionic equation for the reaction occurring. [1]

(ii) Use the standard electrode potentials, $E^{\ominus}$, listed below to explain why chlorine can react with bromide ions but iodine cannot react with bromide ions. [3]

Half-equation	$E^{\ominus}$ / V
$I_2 + 2e^- \rightleftharpoons 2I^-$	+0.54
$Br_2 + 2e^- \rightleftharpoons 2Br^-$	+1.09
$Cl_2 + 2e^- \rightleftharpoons 2Cl^-$	+1.36

(c) Sodium chloride and sodium iodide are both compounds which contain halide ions.

(i) Silver nitrate solution may be used to differentiate between solutions of sodium chloride and sodium iodide. Give the observations that would be expected in **both** cases. [1]

(ii) Both sodium chloride and sodium iodide react with concentrated sulfuric acid. The observations made during both reactions are very different. Discuss the reactions occurring. Your answer should include

- the observations made during both reactions,
- the identities of any products,
- the reasons for any differences in the reactions that occur. [5]

(QWC) [1]

(d) Chlorine produces a range of oxoacids, including chloric(I) acid, HOCl, and chloric(VII) acid, $HClO_4$. Chloric(I) acid is considered to be a weak acid whilst chloric(VII) acid is considered to be a strong acid.

(i) What is meant by the term *strong acid*? [1]

(ii) Write an expression for the acid dissociation constant, K_a, of chloric(I) acid, HOCl. [1]

(iii) The pH of a solution of chloric(I) acid of concentration $0.100 \, mol \, dm^{-3}$ was found to be 4.23. Calculate the concentration of hydrogen ions in this solution. [2]

(iv) Using the information from part (iii), calculate the value of the acid dissociation constant, K_a, for chloric(I) acid. [2]

(v) When the weak acid HOCl reacts with the strong base sodium hydroxide it forms the salt sodium chlorate(I), NaOCl. Suggest a pH value for a solution of NaOCl, giving a reason for your answer. [2]

Total [20]

SECTION A

*Answer **all** questions in the spaces provided.*

1. Potassium peroxodisulfate(VI) (persulfate) is a white crystalline compound of formula $K_2S_2O_8$. It is a powerful oxidising agent and has uses as a food additive, in hair dyes and as a nappy steriliser.

   (a) Unusually for potassium compounds, it is not very soluble in water.

Temperature / °C	Solubility / g per 100 g H_2O
0	1.75
20	5.29

   $1 \, dm^3$ of a saturated solution of potassium persulfate at 20 °C was cooled to 0 °C. Calculate the mass of solid potassium persulfate that crystallised from the solution. [2]

   ....

   ....

   (b) (i) A hot solution of potassium persulfate slowly decomposes, giving oxygen as one of the products.

   $$2K_2S_2O_8(aq) \ + \ 2H_2O(l) \ \longrightarrow \ 4KHSO_4(aq) \ + \ O_2(g)$$

   Calculate the maximum volume of oxygen gas that can be produced at 80 °C when a solution containing 0.100 mol of potassium persulfate decomposes as shown above. [2]

   [At 80 °C 1 mol of oxygen has a volume of $29.0 \, dm^3$]

   ....

   ....

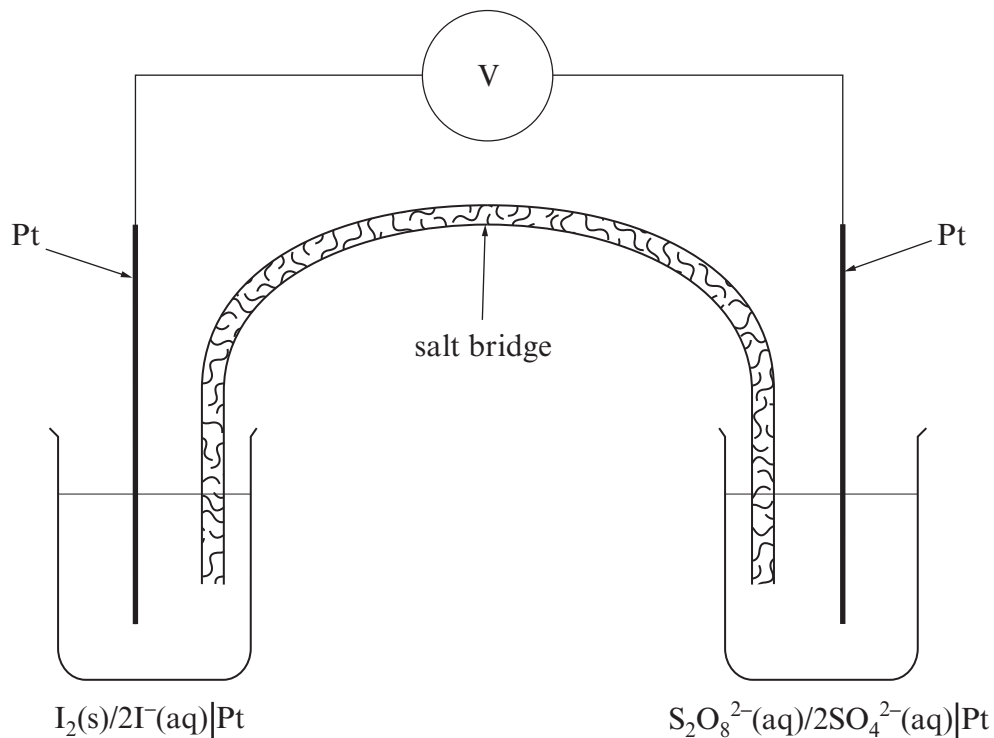   ....

   (ii) Suggest a way that the rate of decomposition of the potassium persulfate solution described in (i) could be measured. [1]

   ....

   ....

   ....

(c)　　The diagram below shows a cell that uses persulfate ions in aqueous solution.

$I_2(s)/2I^-(aq)|Pt$　　　　　　　　　　　　　　$S_2O_8{}^{2-}(aq)/2SO_4{}^{2-}(aq)|Pt$

(i)　　State the role of the platinum electrodes in this cell.　　　　　　　　　　[1]

....

(ii)　　Use the information given in the equations to state and explain the direction of electron flow in the external circuit.　　　　　　　　　　　　　　　[2]

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

$$I_2(s)\qquad\ +\ 2e^-\ \rightleftharpoons\ 2I^-(aq)\qquad\quad E^{\ominus} = +0.54\ V$$

....

....

....

(c) Sodium chloride and sodium iodide both react with concentrated sulfuric acid to give the corresponding hydrogen halide e.g.

$$NaI \quad + \quad H_2SO_4 \quad \longrightarrow \quad NaHSO_4 \quad + \quad HI$$
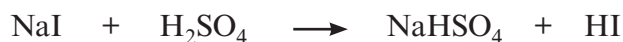
However, the reaction with sodium iodide continues, giving hydrogen sulfide and iodine as two of the products. This further type of reaction does not occur when sodium chloride is used in place of sodium iodide.

(i) Describe what is **seen** when solid sodium iodide is added to concentrated sulfuric acid. [2]

....

....

....

(ii) The following equations show the standard electrode potentials for the Cl_2/Cl^- and I_2/I^- systems.

$$Cl_2 \quad + \quad 2e^- \quad \rightleftharpoons \quad 2Cl^- \qquad E^\ominus = +1.36\,V$$

$$I_2 \quad + \quad 2e^- \quad \rightleftharpoons \quad 2I^- \qquad E^\ominus = +0.54\,V$$

Use these values to explain why only hydrogen iodide (represented as I^- in the equation) is able to further react with concentrated sulfuric acid in this way. [2]

....

....

....

....

(d) The reaction of chlorine with sodium hydroxide solution gives aqueous sodium chlorate(I) as one of the chlorine-containing products.

(i) Give the equation for this reaction. [1]

....

(ii) State **one** use for a solution of sodium chlorate(I). [1]

....

Total [11]

SECTION B

*Answer **both** questions in the separate answer book provided.*

4. *(a)* In the reaction below carbon monoxide is acting as a reducing agent.

$$5CO \ + \ I_2O_5 \ \longrightarrow \ 5CO_2 \ + \ I_2$$

Use oxidation states (numbers) to show that carbon monoxide is acting as a reducing agent in this reaction. [2]

(b) State how the stabilities of the +II and +IV oxidation states vary down Group 4. [1]

(c) You are given two solutions. One contains aqueous aluminium ions, Al^{3+}, and the other contains aqueous lead(II) ions, Pb^{2+}.

(i) Describe a reaction to show that both of these ions exhibit amphoteric behaviour. Your answer should state the reagent(s) used, the names of any precipitates and any relevant observations. *Chemical equations are not required.* [4]
QWC [1]

(ii) Describe what is **seen** when iodide ions are added to an aqueous solution of Pb^{2+} ions. Give the **ionic** equation for the reaction that occurs. [2]

(d) Monomeric aluminium chloride is described as containing an electron-deficient species.

(i) Explain, using monomeric covalent aluminium chloride, what is meant by *electron deficient* and why this leads to the ready formation of the Al_2Cl_6 dimer. You should show the structure of this dimer as part of your answer. [3]

(ii) The electron-deficient nature of the aluminium chloride monomer results in the compound having an affinity for chlorine-containing species. This is important in catalysis and also in the production of specialised solvents. Give **one** example of the use of the monomer in either of these ways. [1]

(iii) On heating, gaseous dimeric aluminium chloride molecules dissociate into the monomer.
$$Al_2Cl_6(g) \ \rightleftharpoons \ 2AlCl_3(g)$$
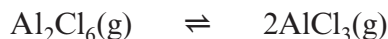
I State **one** reason why the entropy of this gaseous system is increasing. [1]

II Use the equation

$$\Delta G = \ \Delta H \ - \ T\Delta S$$

to calculate the temperature at which the dissociation of gaseous Al_2Cl_6 molecules into gaseous $AlCl_3$ molecules just occurs spontaneously. The entropy change for this reaction, ΔS, is 88 J mol^{-1} K^{-1} and the enthalpy change, ΔH, is 60 kJ mol^{-1}. [2]

SECTION B

*Answer **both** questions in the separate answer book provided.*

4. *(a)* Electrochemical cells are used as power sources in many everyday applications. To decide what to use in a cell, it is necessary to know the standard electrode potential for electrodes. This is measured using a standard hydrogen electrode as a reference standard.

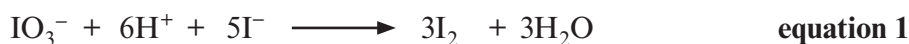Draw a labelled diagram of the apparatus you would use to measure the standard electrode potential of an Fe^{3+}/Fe^{2+} electrode.   [5]

(b) Vanadium is a transition metal that can form compounds with a variety of oxidation states. Zinc however forms compounds with an oxidation state of +2 only.

(i) Why can transition elements form compounds with a variety of oxidation states?   [1]

(ii) Give the electronic structure of Zn.   [1]

(iii) State why zinc forms Zn^{2+}.   [1]

You will need the standard electrode potentials in the table below to answer part *(c)*.

| Oxidation state of vanadium at start of reaction	Reaction	$E^{\ominus}/V$
+5	$VO_3^-(aq) + 4H^+(aq) + e \rightleftharpoons VO^{2+}(aq) + 2H_2O(l)$	+1.00
+4	$VO^{2+}(aq) + 2H^+(aq) + e \rightleftharpoons V^{3+}(aq) + H_2O(l)$	+0.34
+3	$V^{3+}(aq) + e \rightleftharpoons V^{2+}(aq)$	−0.26
+2	$V^{2+}(aq) + 2e \rightleftharpoons V(s)$	−1.13
	$Zn^{2+}(aq) + 2e \rightleftharpoons Zn(s)$	−0.76
	$Cu^{2+}(aq) + 2e \rightleftharpoons Cu(s)$	+0.34

(c) Vanadium(V)(aq), as VO_3^-, is yellow and can be reduced by zinc and aqueous acid producing a series of coloured solutions until the reduction stops with the formation of a violet solution. The reducing agent involves the $Zn^{2+}(aq)/Zn(s)$ equilibrium.

(i) State the identity of the violet vanadium-containing solution produced in this reduction. Use standard electrode potentials to explain your answer. [3]

(ii) What is the standard potential of a cell formed from a standard $Zn^{2+}(aq)/Zn(s)$ electrode and a standard $Cu^{2+}(aq)/Cu(s)$ electrode? [1]

(iii) Write the equilibrium equation for the change occurring at the zinc electrode showing the direction in which the reaction proceeds. [1]

(iv) Use Le Chatelier's principle to predict the effect on the electrode potential of the zinc electrode of increasing the concentration of $Zn^{2+}(aq)$ in the electrode. Explain your answer. [2]

(d) Halogens can also form compounds with a variety of oxidation states. Some of these including compounds of iodate(V), IO_3^-, behave as oxidising agents.

A student was investigating the reaction that occurs when iodate(V) oxidises iodide ions to produce iodine. Two possible equations were suggested.

$$IO_3^- + 6H^+ + 5I^- \longrightarrow 3I_2 + 3H_2O \qquad \textbf{equation 1}$$

$$IO_3^- + 4H^+ + 4I^- \longrightarrow IO^- + 2H_2O + 2I_2 \qquad \textbf{equation 2}$$

He prepared a solution of potassium iodate(V) by dissolving 0.978 g of KIO_3 in 250 cm³ of solution. He pipetted 25.0 cm³ of this solution into a conical flask, added excess potassium iodide and titrated the iodine produced with 0.100 mol dm⁻³ sodium thiosulfate solution, $Na_2S_2O_3$. A volume of 27.40 cm³ of this solution was needed to react with the iodate(V).

The equation for the reaction of thiosulfate with iodine is shown below.

$$2S_2O_3^{2-} + I_2 \longrightarrow S_4O_6^{2-} + 2I^-$$
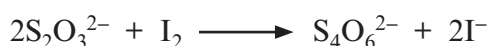
(i) Calculate the number of moles of thiosulfate used to react with the iodine. [1]

(ii) Deduce the number of moles of iodine present in the 25.0 cm³ sample. [1]

(iii) Calculate the number of moles of KIO_3 present in 250 cm³ of the original solution and hence the number of moles present in 25.0 cm³. [1]

(iv) Use your results from (ii) and (iii) to deduce which of **equation 1** and **equation 2** suggested above, correctly shows what happens when iodate(V) ions oxidise iodide ions. Show, by calculation, how you came to this conclusion. [2]

Total [20]

3. *Read the passage below and then answer the questions in the spaces provided.*

Hydrogen Peroxide

If a non-scientist knows only one chemical formula it is most likely to be H_2O for water but how much do you know about another hydrogen oxide, hydrogen peroxide? A molecule of hydrogen peroxide has the molecular formula H_2O_2.

5 Most chemistry students first meet hydrogen peroxide as a colourless solution that is used to prepare oxygen. Bottles of hydrogen peroxide from a pharmacist are often labelled '20 volume'. This means that one volume of solution decomposes to give 20 volumes of oxygen gas. The equation for the decomposition is:

$$2H_2O_2(aq) \longrightarrow 2H_2O(l) + O_2(g)$$

$$1\,dm^3 \qquad\qquad\qquad 20\,dm^3$$
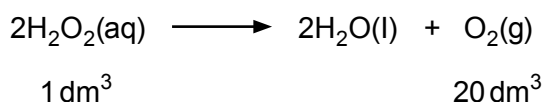
This reaction is very slow at room temperature. However the addition of a suitable catalyst
10 increases the rate of decomposition phenomenally. Manganese(IV) oxide, potatoes and blood are all effective. Potatoes and blood both contain the enzyme catalase and one catalase molecule decomposes 50 000 molecules of H_2O_2 per second!
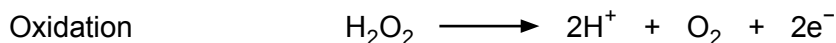
Is hydrogen peroxide an oxidising agent or a reducing agent?

Both in the laboratory and at home hydrogen peroxide is most commonly used as an oxidising
15 agent (so the hydrogen peroxide itself is reduced). The half-equation is:

Reduction $\qquad\qquad H_2O_2 + 2H^+ + 2e^- \longrightarrow 2H_2O$
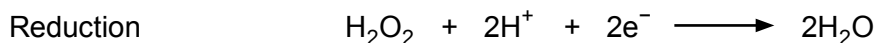
Since some colouring matter is bleached by oxidation and the product of hydrogen peroxide's reduction is water, it is used as a safe bleaching agent particularly in hair treatment. A peroxide blonde is someone with almost white hair, usually as a result of treatment with hydrogen peroxide.

20 However, if hydrogen peroxide reacts with a more powerful oxidising agent such as potassium manganate(VII), the hydrogen peroxide will act as a reducing agent and will itself be oxidised. The half-equation is:

Oxidation $\qquad\qquad H_2O_2 \longrightarrow 2H^+ + O_2 + 2e^-$

Therefore hydrogen peroxide can act as both oxidising agent and reducing agent. In fact, it can
25 react with itself so that alternate molecules are oxidised and reduced. The overall equation is obtained by adding the half-equations for the reduction and oxidation, giving

$$2H_2O_2(aq) \longrightarrow 2H_2O(l) + O_2(g)$$
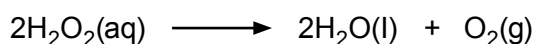
which is the standard decomposition equation!

- End of passage -

(a) Using outer electrons only, draw a dot and cross diagram to show the bonding in a hydrogen peroxide molecule *(line 3)*. [1]

(b) Use the equation for the decomposition of hydrogen peroxide *(line 8)* to calculate the concentration, in mol dm^{-3}, of aqueous hydrogen peroxide solution in a bottle of '20 volume hydrogen peroxide' at 25 °C. [2]

[1 mol of oxygen occupies 24 dm^3 at 25 °C]

Concentration = ... mol dm^{-3}

(c) Manganese(IV) oxide *(line 10)* and potassium manganate(VII) *(lines 20-21)* are typical transition metal compounds.

(i) Give **two** reasons why transition metal compounds can act as catalysts. [2]

....

....

....

(ii) Explain why transition metal complex ions appear coloured. [4]
QWC [1]

....

....

....

....

....

....

....

....

Examiner only

(d) In an acidic solution, hydrogen peroxide is oxidised to oxygen by potassium manganate (VII) *(lines 20-23)*.

(i) Write the half-equation for the reduction of MnO_4^- to Mn^{2+} ions in acidic solution. [1]

....

(ii) Use your answer to (i) and the half-equation given in *line 23* to deduce the overall equation for this reaction. [2]

....

....

....

(iii) $20.0\,cm^3$ of an acidified solution of hydrogen peroxide required $14.80\,cm^3$ of a $0.020\,mol\,dm^{-3}$ solution of potassium manganate(VII) for complete reaction. Calculate the concentration, in $mol\,dm^{-3}$, of the hydrogen peroxide solution. [3]

Concentration = ... $mol\,dm^{-3}$

(e) Explain, using oxidation states, why the decomposition of hydrogen peroxide *(line 27)* can be classified as a redox reaction. [2]

....

....

....

Total [18]

18

Total Section A [40]

(b) Carbon is the first element in Group 4. Two of its allotropes are diamond and graphite. A compound that forms structures corresponding to diamond and graphite is boron nitride.

   (i) Describe the structure of graphite and explain why **hexagonal** boron nitride can adopt the same structure yet have different electrical conductivity properties. [4]

                                                                            QWC [1]

   (ii) State **one** use for the **cubic** boron nitride structure. [1]

(c) Another element in Group 4 is tin. At low temperatures tin exists as its grey form. At higher temperatures the white form is stable. The change can be represented by the equation:

$$Sn_{(grey)} \longrightarrow Sn_{(white)} \qquad \Delta H^{\ominus} = 1.92\,kJ\,mol^{-1}$$

The standard entropy values are $44.8\,J\,K^{-1}\,mol^{-1}$ for grey tin and $51.5\,J\,K^{-1}\,mol^{-1}$ for white tin.

   (i) Calculate the minimum temperature needed to cause grey tin to change to white tin. [3]

   (ii) During Napoleon's disastrous campaign in Russia from June to December in 1812 the tin buttons on his infantry's uniforms disintegrated. Suggest a reason why this might have happened. [1]

(d) An important technological development in recent years has been the hydrogen fuel cell. This uses electrochemical methods to get energy from hydrogen.

   (i) Write the half-equations for the processes occurring at the electrodes and an equation for the overall reaction. [3]

   (ii) Give **one** disadvantage of using hydrogen fuel cells to power vehicles. [1]

Total [20]

5. *(a)* Chlorine reacts with aqueous sodium hydroxide in one of two ways, depending on the temperature used.

   (i) Write the equation for the reaction of chlorine with

        I   cold aqueous sodium hydroxide, [1]

        II   hot aqueous sodium hydroxide. [1]

   (ii) Classify this type of redox reaction. [1]

(b) Chlorine reacts with many elements to form chlorides. Explain why phosphorus forms two chlorides, PCl_3 and PCl_5, but nitrogen only forms NCl_3. [2]

(c) Most ionic chlorides, e.g. sodium chloride, are soluble in water. However some, e.g. silver chloride, are insoluble.

   The enthalpy change of solution of an ionic compound and its solubility depend on the balance between two enthalpy changes. Name these enthalpy changes and state if they are endothermic or exothermic. Explain how the enthalpy change of solution of a compound and its solubility depend on the balance between them. [4]
   QWC [1]

(d) Some standard electrode potentials, $E^{\ominus}$, are given below.

System	$E^{\ominus}$/V
$\frac{1}{2} I_2(s) + e^- \rightleftharpoons I^-(aq)$	+0.54
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$\frac{1}{2} Br_2(l) + e^- \rightleftharpoons Br^-(aq)$	+1.09
$\frac{1}{2} Cl_2(g) + e^- \rightleftharpoons Cl^-(aq)$	+1.36
$Ce^{4+}(aq) + e^- \rightleftharpoons Ce^{3+}(aq)$	+1.45

   (i) Using the information from the table, state which of the **halides** will reduce Fe^{3+} to Fe^{2+}. Give a reason for your answer. [2]

   (ii) Write the cell diagram of the cell formed by combining the $Fe^{3+}(aq)$, $Fe^{2+}(aq)$ and $Ce^{4+}(aq)$, $Ce^{3+}(aq)$ half cells and calculate the standard e.m.f. of this cell. [2]

QUESTION 5 CONTINUES ON PAGE 12