PEARSON EDEXCEL A-LEVEL CHEMISTRY · 9CH0

CHEMISTRY



EXAM QUESTION PACK

TOPIC 3: REDOX I

SCORE	PERCENTAGE
/196	%

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Q1.

In which of the following ions does the metal have an oxidation number of +2?

☒ **A** MnO_4^{2-}

☒ **B** VO^{2+}

☒ **C** $[Fe(CN)_6]^{4-}$

☒ **D** $[CrCl_2(H_2O)_4]^{+}$

(Total for question = 1 mark)

Q2.

Ethanedioic acid, $H_2C_2O_4.\mathbf{x}H_2O$, is a white, hydrated, crystalline solid.

A sample of 0.747 g of this solid is mixed with a small amount of dilute sulfuric acid and dissolved completely in distilled water. The volume is made up to exactly 250 cm^3 in a volumetric flask.

25.0 cm^3 of the resulting solution is titrated with a solution of manganate(VII) ions, $MnO_4^-(aq)$, of concentration 0.0105 mol dm^{-3}. The titration is repeated three more times.

During the titration, the ethanedioic acid is oxidised to form carbon dioxide and the manganate(VII) ions are reduced to Mn^{2+} ions.

(a) The steps followed in the titration are:

Step 1	A burette and pipette are rinsed.
Step 2	The burette is filled with the solution of manganate(VII) ions and the initial volume is recorded.
Step 3	25.0 cm^3 of the ethanedioic acid solution is measured using a pipette and transferred to a conical flask, which is warmed to 60 °C.
Step 4	The solution of manganate(VII) ions is added slowly, whilst swirling the flask.
Step 5	Distilled water is used to rinse any splashes on the inside of the flask.
Step 6	Near the end-point, the solution of manganate(VII) ions is added drop by drop.
Step 7	The volume of the solution of manganate(VII) ions is recorded at the end-point.

(i) State and justify what you would use in **Step 1** to rinse the burette.

(2)

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(ii) Describe how to measure the volume of the solution of manganate(VII) ions accurately in the burette in **Step 2**.

(2)

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(iii) Explain why the use of distilled water in **Step 5** will not affect the results of the titration.

(1)

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(b) Write half-equations for the reduction and oxidation reactions that occur during the titration.

Use these to construct an overall equation for the reaction. State symbols are not required.

(3)

(c) The results for the titration are shown in the table.

Titration	Volume of titre / cm^3
1	23.35
2	22.65
3	23.00
4	22.75

Select the appropriate titres and calculate their mean.

(1)

(d) Calculate the number of moles of water, **x**, in 1 mol of solid ethanedioic acid.

(5)

(e) A student carrying out the titration missed the end-point of the reaction and hence added

too much of the solution of manganate(VII) ions.

Deduce what effect this would have on the student's value of the number of moles of water, **x**.

(2)

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(Total for question = 16 marks)

Q3.

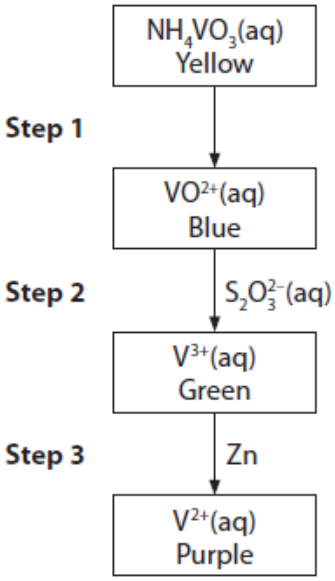
The element vanadium exists in a variety of oxidation numbers in its compounds.

(a) Calculate the oxidation number of vanadium in the compound NH_4VO_3.

(1)

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(b) NH_4VO_3 forms VO_2^+ in solution and can be reduced stepwise to form the ion, $V^{2+}(aq)$.



(i) Use the Data Booklet to give a suitable reducing agent for **Step 1** in order to only produce the blue solution.

(1)

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(ii) Give a possible reason why the reducing agent you choose may not be suitable.

(1)

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(iii) Explain what you would see during **Step 1**.

(2)

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(iv) Justify the choice of reducing agent for **Step 2**.
Write an equation for the reaction. State symbols are not required.

(2)

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Q4.

Titanium is the seventh most abundant metal in the Earth's crust and occurs principally as rutile (impure titanium(IV) oxide, TiO_2).

Early attempts to extract the metal from its oxide by reduction with heated carbon failed because the compound titanium carbide is formed. In 1910, however, pure titanium was made by heating titanium(IV) chloride with sodium.

Titanium has a high melting temperature and a density of 4.50 g cm^{-3}. Titanium is as strong as steel, but is about 40% less dense and is therefore suitable for use in the aircraft industry.

Titanium metal resists corrosion as it has an impervious coating of titanium(IV) oxide. The metal adheres well to bone, is not rejected by the body and is in demand for the manufacture of replacement joints.

Titanium has two common oxidation states, +3 and +4. Solutions containing the $[Ti(H_2O)_6]^{3+}(aq)$ ion, such as titanium(III) chloride, are purple in colour and are readily oxidized by the oxygen in air to colourless titanium(IV) ions. An aqueous solution of titanium(III) chloride is a strong reducing agent. Titanium(IV) chloride, $TiCl_4$, is a colourless liquid with a boiling temperature of 136°C. This compound is used, in conjunction with organic compounds of aluminium, as a catalyst for the polymerization of propene to poly(propene). Titanium(IV) chloride is hydrolysed by water to give titanium(IV) oxide, TiO_2 and hydrogen chloride gas.

Titanium(IV) oxide is a white, non-toxic solid at room temperature. It is used as a white pigment in paint, largely replacing toxic lead compounds which were used previously. Titanium(IV) oxide reacts with concentrated sulfuric acid to form a salt and water. Titanium(IV) oxide also reacts with aqueous potassium hydroxide solution, under suitable conditions, to form a compound with formula $K_2Ti(OH)_6$.

(a) (i) Write the equation for the reaction which occurs during the manufacture of titanium from titanium(IV) chloride as described in the article above. State symbols are not required.

(1)

(ii) Explain, by stating the changes of oxidation numbers, why the reaction in (i) is classified as a redox reaction.

(2)

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(b) Complete the electronic configurations of

(2)

Ti [Ar] .....

Ti^{3+} [Ar] .....

Ti^{4+} [Ar] .....

(c) Use your answer to (b) to explain why titanium is

(i) a *d*-block element

(1)

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(ii) a transition element

(1)

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*(d) (i) Explain why the hexaaquatitanium(III) ion, $[Ti(H_2O)_6]^{3+}$, is coloured.

(3)

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(ii) Explain briefly why titanium(IV) compounds are colourless.

(1)

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*(e) (i) Titanium(IV) oxide has a melting temperature of 1830 °C. Use this data, plus information in the article at the start of the question, to compare the structure and bonding in titanium(IV) oxide with that in titanium(IV) chloride. Hence explain why these two compounds change state at very different temperatures.

(4)

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(ii) Give the term used to describe an oxide, such as titanium(IV) oxide, which can react with both acids and bases.

(1)

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(iii) Using information in the article, write the equation for the reaction between titanium(IV) oxide and aqueous potassium hydroxide solution. State symbols are not required.

(1)

(iv) Titanium(IV) chloride is one of the catalysts used in the polymerization of propene to form poly(propene).

Give the displayed formula of the repeat unit of poly(propene).

(1)

(f) The concentration of a solution of titanium(III) chloride can be determined by titration with a solution of hydrogen peroxide, H_2O_2, in acidic conditions. The end-point of the reaction is when the solution of titanium(III) chloride in the flask goes colourless.

(i) Complete the ionic half-equation to show the reduction of hydrogen peroxide.

(1)

$$H_2O_2 + 2H^+ + \text{.....} \rightarrow \text{.....}$$

*(ii) One mole of hydrogen peroxide reacts with two moles of titanium(III) chloride.

In an experiment, 5.00 cm^3 of a sample of titanium(III) chloride solution was transferred to a volumetric flask and made up to 250 cm^3 of an aqueous solution. A 25.0 cm^3 portion of this diluted solution was acidified and titrated with a 0.0200 mol dm^{-3} solution of hydrogen peroxide, H_2O_2. The mean titre was 22.50 cm^3.

Calculate the concentration of the **original** titanium(III) chloride solution, in mol dm^{-3}.

(3)

(iii) Use information in the article to suggest why this titration gives a value that is lower than the true value for the concentration of titanium(III) chloride solutions.

(1)

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(Total for Question = 23 marks)

Q5.

This question is about some redox reactions of chlorine, bromine and iodine.

Chlorine undergoes disproportionation when it reacts with **hot** aqueous sodium hydroxide solution.

(i) Complete the ionic equation for this reaction. State symbols are not required.

(1)

..... Cl_2 + ... OH^- → ..... Cl^- + ... ClO_3^- + ... H_2O

(ii) Explain, in terms of oxidation numbers, why this is a disproportionation reaction.

(2)

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(Total for question = 3 marks)

Q6.

Chlorine is used to prevent the growth of bacteria in swimming pool water. It reacts as shown below.

$$Cl_2(aq) + H_2O(l) \rightleftharpoons HCl(aq) + HClO(aq)$$

(a) (i) By giving appropriate oxidation numbers, explain why this is a disproportionation reaction.

(3)

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(ii) State and explain the effect on the position of equilibrium if concentrated hydrochloric acid is added to a sample of chlorinated swimming pool water.

(2)

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(b) In a similar reaction, chlorine reacts with sodium hydroxide to make household bleach.

$$Cl_2(aq) + 2NaOH(aq) \rightarrow NaCl(aq) + NaClO(aq) + H_2O(l)$$

The concentration of NaClO in diluted bleach was measured by titration.
A 25.0 cm^3 sample of bleach was pipetted into a conical flask. Approximately 1.5 g of solid potassium iodide and 10 cm^3 of hydrochloric acid with concentration 2.00 mol dm^{-3} were added. Each mole of ClO^-, from the NaClO in the solution of bleach, produced one mole of iodine, I_2, which was titrated with sodium thiosulfate solution.

(i) Complete the ionic half-equations below for the reaction of ClO^- with acidified potassium iodide by balancing them and **adding electrons** where required.

(2)

$$ClO^- + \text{.....}H^+ \rightarrow Cl^- + H_2O$$

$$\text{.....}I^- \rightarrow I_2$$

(ii) Use your answer to (a)(i) to write the overall ionic equation for the reaction between ClO^- and I^-ions in acidic conditions.

(1)

(iii) The iodine in the sample required a mean (average) titre of 24.20 cm^3 of 0.0500 mol dm^{-3} sodium thiosulfate solution. Thiosulfate ions react with iodine as shown below.

$$2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$$

Calculate the number of moles of iodine in the solution.

(2)

(iv) What is the number of moles of ClO^- ions in the sample of diluted bleach?

(1)

(v) Hence calculate the concentration, in mol dm^{-3}, of ClO^- in the diluted bleach.

(1)

(vi) 1.5 g of potassium iodide, KI, contains 9.04×10^{-3} mol of I^-. Use your answers to (b)(ii) and (b)(iv) to show by calculation why this amount was suitable.

(2)

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(vii) A student carrying out this titration measured the mean (average) titre as 24.50 cm^{-3}.

What is the percentage difference in this student's titre, compared with the accurate value of 24.20 cm^-?

(1)

(viii) The difference between the student's mean titre and the accurate value was **not** due to the limitations in the accuracy of the measuring instruments.

Suggest one possible reason for this difference.

(1)

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(c) Suggest **one** damaging effect to the upper atmosphere which could be caused by the presence of chlorine compounds.

(1)

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(Total for question = 17 marks)

Q7.

This is a question about chromium(III) and chromium(VI) compounds.

The chromium(III) complex, $[Cr(OH)_6]^{3-}$, can be oxidised to chromate(VI) ions, CrO_4^{2-}

, by hydrogen peroxide solution.

(i) Deduce the oxidation half-equation for this reaction, which takes place in alkaline conditions. State symbols are not required.

(2)

(ii) If the solution of chromate(VI) ions is then acidified, the colour of the solution changes to orange as dichromate(VI) ions form.

Write the equation for this change. State symbols are not required.

(1)

(iii) In acidic conditions, dichromate(VI) ions can also be reduced to chromium(III) ions using hydrogen peroxide.

The value of $E^{\ominus}_{cell}$ cell = + 0.65 V for which the cell diagram is

$$Pt(s) \mid H_2O_2(aq), [2H^+(aq)+O_2(g)] \,\vdots\, [Cr_2O_7^{2-}(aq)+14H^+(aq)], [2Cr^{3+}(aq)+7H_2O(l)] \mid Pt(s)$$

Deduce from the cell diagram the oxidation and the reduction half-equations, and thus the overall equation for this reaction.
State symbols are not required.

(3)

(Total for question = 6 marks)

Q8.

Prussian Blue, $[Fe_4[Fe(CN)_6]_3]$, is a dark blue pigment used in painting and dyeing.

It was discovered around 1700 AD in the German state of Prussia.

Prussian Blue is formed when an iron(III) salt is added to a solution containing the complex ion $[Fe(CN)_6]^{4-}$.

The cyanide ion has the formula CN^-.

(a) (i) The oxidation number of Fe in the $[Fe(CN)_6]^{4-}$ ion is

(1)

☒ **A** +2

☒ **B** +3

☒ **C** +4

☒ **D** +6

(ii) Draw a diagram to show the shape of a $[Fe(CN)_6]^{4-}$ ion, using the structure CN^- to represent a cyanide ligand and showing how the cyanide ligands bond to the central iron ion.

(2)

(b) A solution containing $[Fe(CN)_6]^{3-}$ ions can be made from $[Fe(H_2O)_6]^{2+}$ ions in two steps as shown:

Step 1 $[Fe(H_2O)_6]^{2+}(aq) + 6CN^-(aq) \rightleftharpoons [Fe(CN)_6]^{4-}(aq) + 6H_2O(l)$

Step 2 $2[Fe(CN)_6]^{4-}(aq) + Cl_2(aq) \rightarrow 2[Fe(CN)_6]^{3-}(aq) + 2Cl^-(aq)$

Name the type of reaction taking place in each of Steps 1 and 2.

(2)

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(c) In a separate reaction, aqueous sodium hydroxide was added to a solution containing iron(II) sulfate. A green precipitate formed that turned brown on standing in air.

Identify the green precipitate and explain why it turns brown on standing in air.

(3)

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Q9.

Answer the question with a cross in the box you think is correct ☒. If you change your mind about an answer, put a line through the box ☒ and then mark your new answer with a cross ☒.

This question is about the compound potassium bromate, $KBrO_3$.

These bromate ions react with bromide ions in acidic solution.

$$BrO_3^-(aq) + 5Br^-(aq) + 6H^+(aq) \rightarrow 3Br_2(aq) + 3H_2O(l)$$

(i) Explain, in terms of oxidation numbers, whether or not this is a disproportionation reaction.

(2)

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(ii) What is the overall order of this reaction?

(1)

☒ **A** 3

☒ **B** 6

☒ **C** 12

☒ **D** cannot tell from this information

(Total for question = 3 marks)

Q10.

Persulfate ions, $S_2O_8^{2-}$, oxidize iodide ions in aqueous solution to form iodine and sulfate ions, SO_4^{2-}.

(a) Write the ionic equation for this reaction. State symbols are not required.

(1)

(b) The effect of persulfate ion concentration on the rate of this reaction was measured.

A few drops of starch solution and a small measured volume of sodium thiosulfate solution were added to the potassium persulfate solution.

Potassium iodide solution was then added and the time taken for the mixture to change colour was measured.

The reaction was repeated using different concentrations of potassium persulfate, but the same volumes and concentrations of sodium thiosulfate solution and potassium iodide solution.

The rates of the reaction were compared using the reciprocal of the time (1/time) for the mixture to change colour as a measure of the initial rate.

(i) What is the final colour of the reaction mixture?

(1)

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(ii) What would happen if the reaction was carried out without the addition of sodium thiosulfate?

(1)

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(iii) Explain why the concentration of iodide ions remains constant until the mixture changes colour.

(1)

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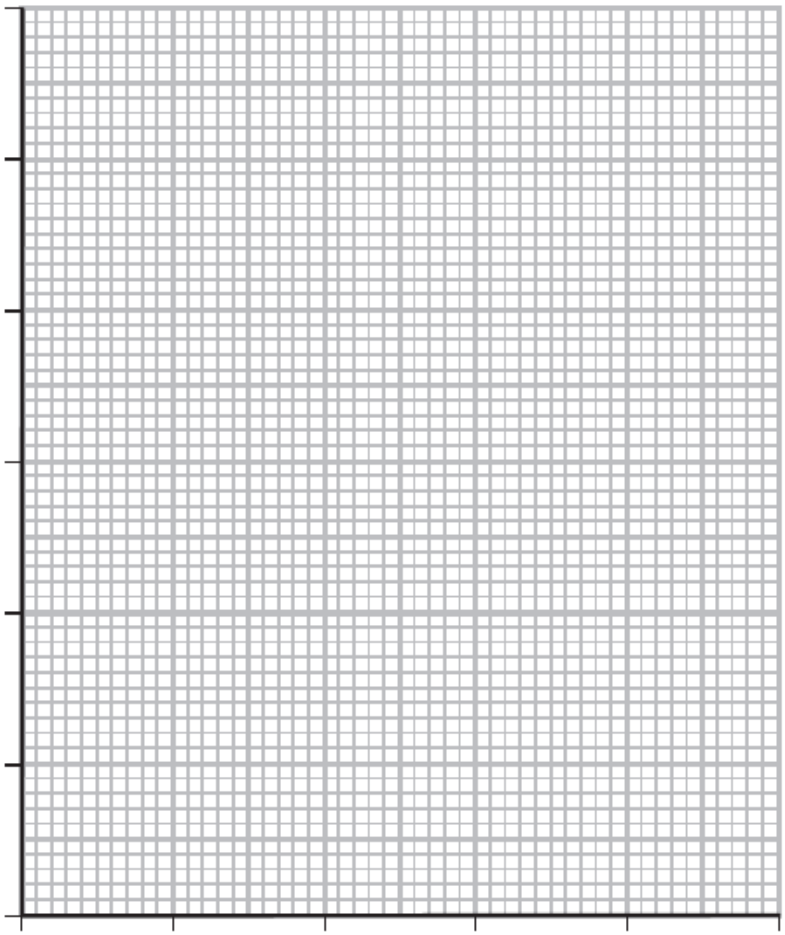
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(c) The results obtained from the experiment in part (b) were tabulated as follows.

$[S_2O_8^{2-}]$ /mol dm^{-3}	Time /s	1/time /s^{-1}
0.0100	40.0	0.0250
0.0090	44.4	0.0225
0.0075	53.3	0.0188
0.0060	66.7	0.0150

(i) Plot a graph of 1/time on the vertical axis against the concentration of the persulfate ions.

(2)



(ii) 1/time is a measure of the initial rate of the reaction.

Deduce the order of the reaction with respect to persulfate ions.

Justify your answer.

(2)

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(iii) The reaction is first order with respect to iodide ions. Write the overall rate equation for the reaction and deduce the units for the rate constant.

(2)

Rate =

Units for the rate constant .....

(Total for question = 10 marks)

Q11.

(i) $Cu^+(aq)$ ions are not stable in solution and undergo a disproportionation reaction.

Suggest an equation for this reaction, including state symbols.

(1)

(ii) Suggest in what way the **appearance** of CuI is similar to that of ZnI_2.

Give a reason for this similarity.

(2)

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Q12. Chlorine is prepared in the laboratory by the reaction of potassium manganate(VII) with concentrated hydrochloric acid. The ionic half-equations and their standard electrode potentials are

$$\tfrac{1}{2}Cl_2(g) + e^- \rightleftharpoons Cl^-(aq) \qquad E^\ominus = +1.36\ V$$

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l) \qquad E^\ominus = +1.51\ V$$

(a) Concentrated hydrochloric acid is preferred to 1 mol dm^{-3} hydrochloric acid because

☒ **A** it lowers the activation energy of the reaction.

☒ **B** the reaction is thermodynamically unfavourable under standard conditions.

☒ **C** it increases the rate of reaction.

☒ **D** it increases the equilibrium constant, K_c, for the overall reaction.

(b) From the half-equations it may be deduced that

☒ **A** 1 mol of manganate(VII) ions reacts with 5 mol of chloride ions.

☒ **B** 5 mol of manganate(VII) ions reacts with 1 mol of chloride ions.

☒ **C** 1 mol of manganate(VII) ions reacts with 2.5 mol of chlorine.

☒ **D** 2.5 mol of manganate(VII) ions reacts with 1 mol of chlorine.

(Total for Question = 2 marks)

Q13.

Tablets containing potassium manganate(VII), $KMnO_4$, are dissolved in water forming an antiseptic solution to treat skin conditions. The manufacturers claim that each tablet contains 400 mg of $KMnO_4$.

To check the claim, the titration procedure outlined was carried out.

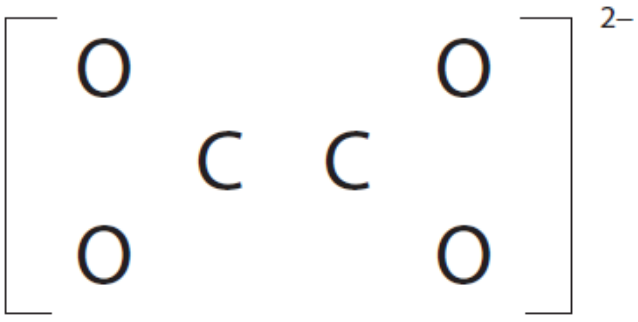
- Five tablets were dissolved in distilled water to make 100.0 cm^3 of solution.
- Some of the $KMnO_4$ solution was used to fill a burette.
- 25.0 cm^3 of sodium ethanedioate solution, $Na_2C_2O_4(aq)$, of concentration 0.200 mol dm^{-3}, was added to a conical flask and warmed.
- Sulfuric acid, of concentration 2 mol dm^{-3}, was also added to the conical flask.
- The $KMnO_4$ solution was added to the flask from the burette, until the end-point.

The equation for the reaction between MnO_4^- ions from the $KMnO_4$ and $C_2O_4^{2-}$ ions from the sodium ethanedioate solution is shown.

$$16H^+(aq) + 2MnO_4^-(aq) + 5C_2O_4^{2-}(aq) \rightarrow 2Mn^{2+}(aq) + 10CO_2(g) + 8H_2O(l)$$

(i) Complete the dot-and-cross diagram for the ethanedioate ion.
Show the outer electrons only.

(2)



(ii) Determine the oxidation number of carbon in the ethanedioate ion, $C_2O_4^{2-}$.

(1)

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(Total for question = 3 marks)

Q14.

The oxidation state of nickel is **not** +2 in

☒ **A** $[Ni(CO)_4]$

☒ **B** $[Ni(H_2O)_4(OH)_2]$

☒ **C** $[Ni(NH_3)_6]^{2+}$

☒ **D** $[Ni(CN)_4]^{2-}$

(Total for question = 1 mark)

Q15.

Which of the following is a redox reaction?

☒ **A** $Cr_2O_7^{2-} + 2OH^- \rightarrow 2CrO_4^{2-} + H_2O$

☒ **B** $[Cu(H_2O)_6]^{2+} + 4Cl^- \rightarrow [CuCl_4]^{2-} + 6H_2O$

☒ **C** $4OH^- + 4MnO_4^- \rightarrow 4MnO_4^{2-} + 2H_2O + O_2$

☒ **D** $[Fe(H_2O)_6]^{3+} + 3OH^- \rightarrow [Fe(H_2O)_3(OH)_3] + 3H_2O$

(Total for question = 1 mark)

Q16.

This question is about redox reactions.

Name the ion with formula PO_3^{3-} . Include the relevant oxidation number.

(1)

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(Total for question = 1 mark)

Q17.

This question is about redox reactions.

Manganese(IV) oxide, MnO_2 , and manganate(VII) ions, MnO_4^- , react in alkaline solution to form manganate(VI) ions, MnO_4^{2-} .

(i) Write the **ionic** equation for this reaction.

State symbols are not required.

(2)

(ii) Give a reason why this reaction is **not** disproportionation.

(1)

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(Total for question = 3 marks)

Q18.

Benzene

Benzene was first discovered by Michael Faraday in London in 1825. He identified it by distilling

a liquid obtained from condensing a gas produced by heating whale oil.

In 1834, Mitscherlich discovered the same liquid could be obtained by heating benzoic acid with lime.

Benzene is now produced from petroleum. Petroleum is fractionally distilled and benzene is obtained from the hexane in the naphtha fraction.

Hexane is heated to about 770 K. It then passes to a reactor where it reacts to form cyclohexane and hydrogen. The cyclohexane is then dehydrogenated to form benzene. Other aromatic products like methylbenzene and dimethylbenzenes are also produced. The aromatic products are separated by further distillation.

Benzene is the starting compound for a large number of useful chemicals and materials. For example it is used to make the polymer, poly(phenylethene).

(a) Write an equation for the reaction between benzoic acid and lime, calcium oxide. State symbols are not required.

(1)

(b) Deduce, with references to oxidation numbers, what has happened to the carbon in the conversion of cyclohexane to benzene.

(2)

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Q19.

This question is about chlorine.

Write the formula of potassium chlorate(V).

(1)

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(Total for question = 1 mark)

Q20. In nitric(V) acid, HNO_3, the oxidation number of the nitrogen is +5. This means that the **nitrogen** in nitric acid

☒ **A** has five electrons in its outer shell.

☒ **B** is an ion with a charge of +5.

☒ **C** would have a charge of +5 if its bonding electrons were transferred completely.

☒ **D** forms five covalent bonds in total.

(Total for Question = 1 mark)

Q21.

This question is about transition metals.

In which of these complex ions does the transition metal have the oxidation number +3?

(1)

☒ **A** $[Ag(CN)_2]^-$

☒ **B** $[CuCl_4]^{2-}$

☒ **C** $[Fe(CN)_6]^{3-}$

☒ **D** $[Ni(EDTA)]^{2-}$

(Total for question = 1 mark)

Q22.

Answer the question with a cross in the box you think is correct ☒ . If you change your mind about an answer, put a line through the box ☒ and then mark your new answer with a cross ☒ .

The halogens are elements in Group 7 of the Periodic Table.

Sodium chlorate(I) is a bleaching agent.

(i) Sodium chlorate(I) can be made by the reaction of chlorine with sodium hydroxide.

Show, by using oxidation numbers, that this reaction is disproportionation.

$$2NaOH + Cl_2 \rightarrow NaClO + NaCl + H_2O$$

(2)

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(ii) A different bleaching agent can be made by the reaction of chlorine with sodium hydroxide under different conditions.

Balance this equation.

$$.....NaOH + ...Cl_2 \rightarrow NaClO_3 + ...NaCl + ...H_2O$$

(1)

(iii) What conditions are required for the reaction in (b)(ii)?

(1)

☒ **A** cold and dilute alkali

☒ **B** cold and concentrated alkali

☒ **C** hot alkali

☒ **D** excess chlorine

(Total for question = 4 marks)

Q23.

In which of the following ions does the metal have an oxidation number of +3?

☒ **A** MnO_4^{2-}

☒ **B** VO^{2+}

☒ **C** $[Fe(CN)_6]^{4-}$

☒ **D** $[CrCl_2(H_2O)_4]^{+}$

(Total for question = 1 mark)

Q24.

This question is about the properties of transition elements, their ions and their complexes.

Give the oxidation state of vanadium in the compound NH_4VO_3.

(1)

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(Total for question = 1 mark)

Q25.

Iron is a transition element that forms a number of ions with the iron in different oxidation states.

(a) What is the electronic configuration of the iron cation that can form the complex ion $[Fe(CN)_6]^{4-}$?

(1)

☒ **A** [Ar] $3d^4 4s^2$

☒ **B** [Ar] $3d^5 4s^0$

☒ **C** [Ar] $3d^6 4s^0$

☒ **D** [Ar] $3d^6 4s^2$

(b) Iron forms a ferrate ion of formula FeO_4^{2-}.

(i) What is the oxidation number of the iron in FeO_4^{2-}?

(1)

☒ **A** + 2

☒ **B** + 3

☒ **C** + 4

☒ **D** + 6

(ii) Deduce, using the data shown, whether ferrate ions are stable in acidic conditions.

Write an overall ionic equation for any reaction that takes place. State symbols are not required.

$$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l) \qquad E^\ominus = +1.23\,V$$

$$FeO_4^{2-}(aq) + 8H^+(aq) + 3e^- \rightleftharpoons Fe^{3+}(aq) + 4H_2O(l) \qquad E^\ominus = +2.20\,V$$

(4)

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(c) The salt $K_3Fe(CN)_6$ can be prepared by oxidising $K_4Fe(CN)_6$.
The relevant standard electrode (redox) potential is:

$$[Fe(CN)_6]^{3-}(aq) + e^- \rightleftharpoons [Fe(CN)_6]^{4-}(aq) \qquad E^\ominus = +0.36\,V$$

Which species will oxidise $K_4Fe(CN)_6$ to $K_3Fe(CN)_6$?
You may find it helpful to refer to the Data booklet.

(1)

☒ **A** Ag(s)

☒ **B** $Cl_2(g)$

☒ **C** $Cu^{2+}(aq)$

☒ **D** $Fe^{2+}(aq)$

Q26.

This question is about halogens and redox reactions.

Use these electrode potentials to answer the following questions.

Electrode reaction	$E^{\ominus}$ / V
$I_2(aq) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$Br_2(aq) + 2e^- \rightleftharpoons 2Br^-(aq)$	+1.09
$MnO_2(s) + 4H^+(aq) + 2e^- \rightleftharpoons Mn^{2+}(aq) + 2H_2O(l)$	+1.23
$Cl_2(aq) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51

(i) Which species will oxidise $Fe^{2+}(aq)$ to $Fe^{3+}(aq)$?

(1)

☒ **A** $Br_2(aq)$

☒ **B** $Cl^-(aq)$

☒ **C** $I_2(aq)$

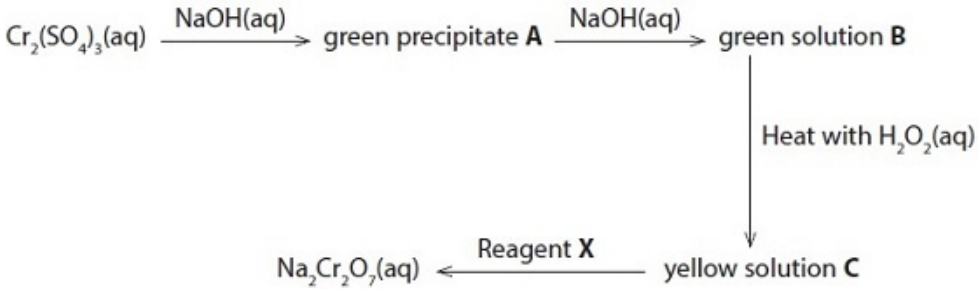
☒ **D** $Mn^{2+}(aq)$

(ii) Write the ionic equation and calculate the $E^{\ominus}_{cell}$ for the reaction between MnO_4^- ions and Br^- ions in acidic solution.
State symbols are not required.

(3)

(Total for question = 4 marks)

Q27. Sodium dichromate(VI) may be prepared from chromium(III) sulfate using the sequence outlined below.



(a) (i) Complete the table below, giving the **formula** of the chromium containing species.

(4)

Substance	Formula of chromium containing species
$Cr_2(SO_4)_3(aq)$	
Green precipitate **A**	
Green solution **B**	
Yellow solution **C**	

(ii) Complete the half equation for the reduction of hydrogen peroxide, $H_2O_2(aq)$.

(1)

$$H_2O_2 + \text{.....} \rightarrow \text{.....} OH^-$$

(iii) Identify reagent **X**

(1)

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(iv) Write an **ionic** equation for the conversion of solution **C** to $Na_2Cr_2O_7$. State symbols are not required.

(1)

(b) Chromium also exists in a +2 oxidation state. Use the Data booklet to determine the feasibility of the disproportionation of chromium(III) in aqueous solution into chromium(II) and dichromate(VI).

In your answer, show both half equations and the overall equation for the proposed disproportionation. Calculate $E^{\ominus}_{cell}$ for this reaction and use this value to predict the feasibility of this reaction.

(4)

(Total for Question = 11 marks)

Q28.

Chlorine and bromine are elements in Group 7 of the Periodic Table.

Both elements exist in a number of different oxidation numbers and therefore are involved in many redox reactions.

(a) Write an equation for the reaction between chlorine and cold, dilute aqueous sodium

hydroxide. State symbols are not required.

(1)

(b) Chlorine dioxide reacts with cold, dilute aqueous sodium hydroxide.

The ionic equation for the reaction is:

$$2ClO_2(aq) + 2OH^-(aq) \rightarrow ClO_2^-(aq) + ClO_3^-(aq) + H_2O(l)$$

Using oxidation numbers, explain why the chlorine in ClO_2 has undergone disproportionation.

(3)

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(c) Chloride ions (Cl^-) can be oxidised to chlorine molecules (Cl_2) by manganate(VII) ions.

A dilute solution containing manganate(VII) ions (MnO_4^-) and an excess of dilute sulfuric acid is added to the solution containing chloride ions.

As the manganate(VII) ion solution is added, it changes from purple to colourless.

(i) The formula of the manganese species that is formed during this reaction is

(1)

☒ **A** $Mn^{2+}(aq)$

☒ **B** $Mn^{3+}(aq)$

☒ **C** $Mn^{4+}(aq)$

☒ **D** $MnO_4^{2-}(aq)$

(ii) Using oxidation numbers, deduce the molar ratio of MnO_4^- to Cl^- that would appear in the balanced chemical equation for the reaction.

(2)

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(d) Potassium bromide reacts with concentrated sulfuric acid. Three of the products of the reaction are hydrogen bromide, bromine and sulfur dioxide.

(i) Explain why the hydrogen bromide, which is a colourless gas, appears as misty fumes when

it makes contact with moist air.

(2)

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(ii) State the type of reaction that occurs in the formation of sulfur dioxide.

(1)

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Q29.

Many vehicles are fitted with airbags which provide a gas-filled safety cushion to protect the occupant of the vehicle if there is a crash.

(a) The first reaction in airbags is the thermal decomposition of sodium azide, NaN_3, to form sodium and nitrogen gas.

(i) Write the equation for this decomposition of sodium azide.

State symbols are not required.

(1)

(ii) In the reaction in (i), a typical airbag is inflated by about 67 dm^3 of gas.

Calculate the **minimum mass** of sodium azide, in grams, needed to produce this volume of gas. Use the Ideal Gas Equation and give your answer to an appropriate number of significant figures.

For the purpose of this calculation, assume that the temperature is 300 °C and the pressure is 140 000 Pa.

(4)

(b) The second reaction in the airbag is between the sodium produced in the reaction (a)(i) and potassium nitrate.

$$\text{.....}Na + \text{.....}KNO_3 \rightarrow \text{.....}K_2O + \text{.....}Na_2O + \text{.....}N_2$$

Balance the above equation, justifying your answer in terms of the changes in oxidation numbers.

(3)

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(c) The third reaction in the airbag is between the metal oxides and silicon dioxide.

State the type of reaction taking place and justify why this reaction is necessary.

(3)

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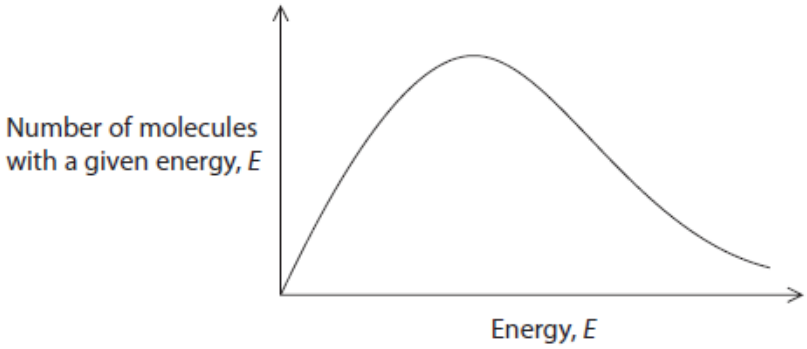
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(d) The Maxwell-Boltzmann distribution diagram shows the molecular energies for the gaseous system immediately after the airbag has been deployed.



What is the change in shape of the curve when the airbag **cools**?

(1)

☒ **A** the peak would shift to the left and be higher

☒ **B** the peak would shift to the left and be lower

☒ **C** the peak would shift to the right and be higher

☒ **D** the peak would shift to the right and be lower

(Total for question = 12 marks)

Q30.

Magnesium nitrate decomposes on heating as shown by the equation.

$$2Mg(NO_3)_2 \rightarrow 2MgO + 4NO_2 + O_2$$

Explain, in terms of all the relevant oxidation numbers, why this is a redox reaction.

(3)

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(Total for question = 3 marks)

Q31.

The reaction between iron and nickel(II) sulfate may be represented by the ionic equation

(a) This reaction is classified as a redox reaction because

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

(1)

☒ **A** the nickel(II) ions are oxidized and iron acts as an oxidizing agent.

☒ **B** the nickel(II) ions are oxidized and iron acts as a reducing agent.

☒ **C** the nickel(II) ions are reduced and iron acts as a reducing agent.

☒ **D** the nickel(II) ions are reduced and iron acts as an oxidizing agent.

(b) This reaction proceeds spontaneously, therefore

(1)

☒ **A** E_{cell} and ΔS_{total} for this reaction must both be positive.

☒ **B** E_{cell} and ΔS_{total} for this reaction must both be negative.

☒ **C** E_{cell} for this reaction must both be positive and ΔS_{total} negative.

☒ **D** E_{cell} for this reaction must both be negative and ΔS_{total} positive.

(Total for question = 2 marks)

Q32.

This question is about redox reactions.

State what happens to a reducing agent during a reaction, in terms of oxidation number **and** electrons.

(1)

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(Total for question = 1 mark)

Q33.

Chlorine compounds take part in many redox reactions.

(a) State, in terms of electron transfer, what is meant by the term oxidising agent.

(1)

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(b) Chlorine dioxide gas, ClO_2, is used to whiten flour. It is unstable and hence is prepared in situ by the reaction:

$$2ClO_3^-(aq) + SO_2(g) + H^+(aq) \rightarrow 2ClO_2(g) + HSO_4^-(aq)$$

Explain, using oxidation numbers, which species is acting as the oxidising agent in this reaction.

(2)

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(c) Chlorine dioxide reacts in alkaline solution to form a mixture of ClO_2^- and ClO_3^- ions.

(i) Write an equation for this reaction. State symbols are not required.

(2)

(ii) Give the reason why this reaction is disproportionation.

(1)

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Q34.

This question is about transition metals and transition metal complexes.

Aqueous vanadium(II) chloride, $VCl_2(aq)$, can be oxidised by bubbling gaseous chlorine, $Cl_2(g)$, through the solution in the absence of air.

40.0 cm^3 of 0.100 mol dm^{-3} VCl_2 solution was oxidised by 144 cm^3 of chlorine gas, at room temperature and pressure (r.t.p.).
The chlorine was reduced to chloride ions, according to the half-equation

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

[Molar volume of a gas at r.t.p. = 24.0 dm^3 mol^{-1}]

(i) Use these data to calculate the final oxidation state of vanadium.
You **must** show your working.

(5)

(ii) State the initial and final colours you would see as the chlorine bubbles through the aqueous vanadium(II) chloride, $VCl_2(aq)$.

(2)

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(Total for question = 7 marks)

Q35.

Copper metal is oxidized to Cu^{2+} by nitrate(V) ions which are reduced to nitrogen monoxide, NO.

By considering the changes to the oxidation numbers of copper and nitrogen, it can be deduced that in this reaction

☒ **A** 2 mol of copper reacts with 3 mol of nitrate(V) ions.

☒ **B** 2 mol of copper reacts with 5 mol of nitrate(V) ions.

☒ **C** 3 mol of copper reacts with 2 mol of nitrate(V) ions.

☒ **D** 5 mol of copper reacts with 2 mol of nitrate(V) ions.

(Total for question = 1 mark)

Q36.

Sulfur dioxide reacts with hydrogen sulfide to form water and sulfur. By considering the changes in the oxidation numbers of sulfur, it can be deduced that, in this reaction

☒ **A** 1 mol of sulfur dioxide oxidizes 2 mol of hydrogen sulfide.

☒ **B** 1 mol of sulfur dioxide reduces 2 mol of hydrogen sulfide.

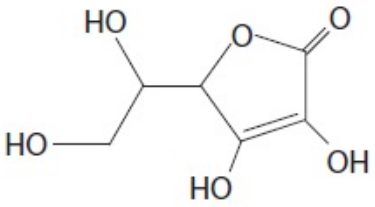
☒ **C** 2 mol of sulfur dioxide oxidizes 1 mol of hydrogen sulfide.

☒ **D** 2 mol of sulfur dioxide reduces 1 mol of hydrogen sulfide.

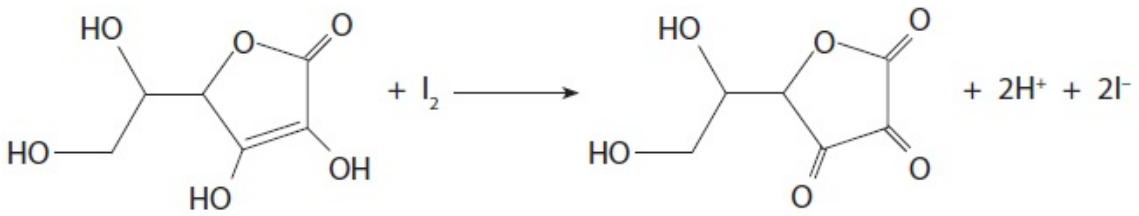
(Total for question = 1 mark)

Q37.

Vitamin C (L-ascorbic acid) is present in fresh fruit and vegetables although prolonged cooking destroys it. The structure of ascorbic acid, $C_6H_8O_6$, is shown below.



(a) The amount of ascorbic acid present in a sample is determined by reacting it with a known amount of iodine. The excess iodine is then measured by titration with a solution of sodium thiosulfate, using a starch indicator. The equations for the reactions are given below.



$$2S_2O_3^{2-} + I_2 \longrightarrow S_4O_6^{2-} + 2I^-$$

Four 500 mg vitamin C tablets were dissolved in distilled water and the solution made up to 250 cm^3 in a volumetric flask. A 25.0 cm^3 portion of this solution was added to an iodine solution containing 2.00×10^{-3} moles of iodine. The resulting mixture was titrated with sodium thiosulfate solution of concentration 0.0631 mol dm^{-3}. The titration was repeated and the mean (average) titre was 27.85 cm^3.

(i) State the type of chemical change undergone by ascorbic acid when it reacts with iodine.

(1)

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(ii) State the colour of the solution in the flask just before the starch is added to the titration mixture and state the colour change at the end-point of the titration.

(2)

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(iii) Calculate the percentage by mass of ascorbic acid in the tablets. The molar mass of ascorbic acid is 176 g mol^{-1}.

(5)

(iv) Explain why using four tablets in 250 cm^3 of solution gives a more accurate result than two tablets in 250 $cm^{3.}$

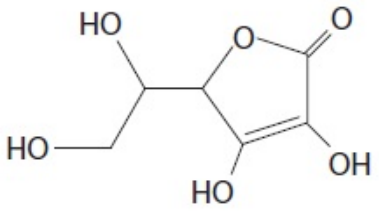
(1)

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(b) The structure of ascorbic acid is shown again below. Vitamin C is one of the optical isomers of this structure.



(i) Mark on this diagram the two chiral centres of this molecule.

(2)

(ii) How might you show that vitamin C is a single optical isomer and not a racemic mixture of the optical isomers of ascorbic acid?

(2)

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(iii) Despite its name, ascorbic acid is not a carboxylic acid but it does contain an ester group. Suggest what happens to destroy vitamin C on prolonged cooking.

(1)

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