PEARSON EDEXCEL A-LEVEL CHEMISTRY · 9CH0

CHEMISTRY



EXAM QUESTION PACK

TOPIC 14: REDOX II

SCORE	PERCENTAGE
/369	%

STEMUP

These questions belong to Pearson Edexcel, not to us.

The examination questions and mark schemes in this pack were written and published by Pearson Edexcel. They are past assessment materials, reproduced here so that students can revise from them.

What we do not own

STEMUP is not affiliated with Pearson Edexcel and does not own, and makes no claim to own, any of the question or mark scheme material in this pack. Pearson has not endorsed, approved or been involved in the production of it. Pearson, Edexcel and the qualification titles used here belong to their respective owners and appear for identification only.

Where to find the originals

Pearson publishes its past papers and mark schemes free of charge. Always check the current specification at qualifications.pearson.com before you revise, because content and wording can change between series.

If you hold the rights to any material in this pack and would like it amended or removed, contact admin@stemup.co.uk and we will act on it.

Q1. This is a question about Group 2 compounds.

Limewater is a solution of calcium hydroxide, commonly used in the identification of carbon dioxide gas. Since calcium hydroxide is only sparingly soluble in water, technicians often make the solution by adding an excess of the solid calcium hydroxide to the required volume of deionised water, shaking the container and then leaving the mixture to settle. In this way, a saturated solution is produced but it can be of variable concentration.

Two students were each given a sample of limewater, from the same batch, in order to determine its concentration. Using 50.0 cm^3 portions of the limewater, they carried out titrations using 0.100 mol dm^{-3} hydrochloric acid. One of the students obtained the following results:

Titration	Trial	1	2
Final Volume /cm^3	14.50	28.60	42.70
Initial Volume /cm^3	0.00	14.50	28.60
Volume Added /cm^3	14.50	14.10	14.10

The student decided that the mean titre was 14.10 cm^3

The equation for the reaction is:

$$Ca(OH)_2(aq) + 2HCl(aq) \rightarrow CaCl_2(aq) + 2H_2O(l)$$

(a) (i) Calculate the number of moles of hydrochloric acid that reacted.

(1)

(ii) Calculate the number of moles of calcium hydroxide, $Ca(OH)_2$, that reacted with the acid.

(1)

(iii) the concentration of $Ca(OH)_2$, in mol dm^{-3}, in this sample of limewater.

(1)

(iv) Calculate the concentration of $Ca(OH)_2$, in g dm^{-3}, in this sample of limewater. Use the Periodic Table as a source of data.

(2)

(v) This student did not include the trial value when calculating the mean titre. Explain why.

(1)

.....

.....

.....

.....

(vi) The second student obtained a different mean titre value for the experiment and thought that this difference may be due to the use of a faulty pipette.

Suggest a simple method, involving distilled water and a balance, by which the accuracy of the pipette in measuring out exactly 50.0 cm^3 could be checked.

(2)

.....

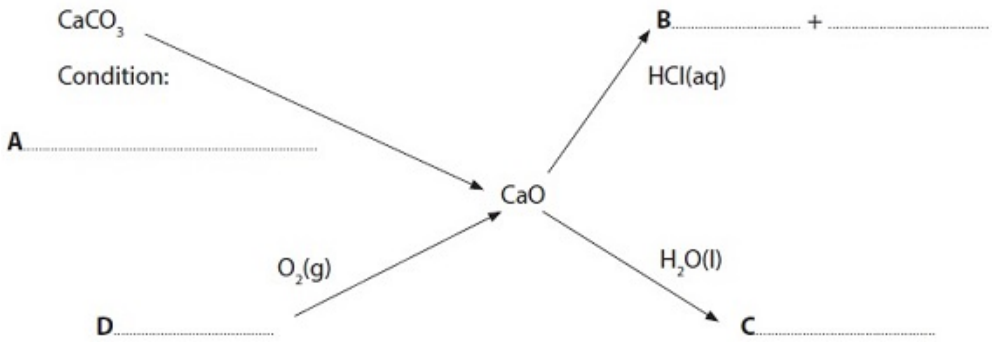
.....

.....

.....

(b) Complete the missing details from the reaction flowchart shown below, giving the condition for **A** and using chemical formulae for answers **B**, **C** and **D**. State symbols are not required.

(4)



(c) In certain areas of the UK, calcium and magnesium carbonates tend to be deposited as an off-white solid on the inside surface of pipes and the surface of heating elements in kettles. These deposits can be removed by treatment with a weak acid. An equation for this is shown below.

$$CaCO_3(s) + 2HA(aq) \rightarrow CaA2(aq) + H_2O(l) + CO_2(g)$$

State **one** observation, other than the solid disappearing, that would be made when the above reaction is carried out.

(1)

.....

.....

(d) The thermal stability of these carbonates depends on a combination of factors, including the size of their lattice energies.

Explain why the lattice energy of calcium carbonate is less exothermic than that of magnesium carbonate.

(2)

.....

.....

.....

.....

(e) Calcium and magnesium ions can be distinguished by the use of a flame test. State the difference in the flame colour and explain how colours in a flame are produced in terms of electronic transitions.

(3)

Calcium

.....

Magnesium

.....

Colour produced by

.....

.....

.....

.....

.....

.....

.....

(Total for Question = 18 marks)

Q2.

A student carried out an investigation to determine the value of ***x*** in hydrated magnesium nitrate(III), $Mg(NO_2)_2 \bullet \boldsymbol{x}H_2O$, using three different methods.

Method 1

- The student prepared an aqueous solution by dissolving 1.15 g of $Mg(NO_2)_2 \bullet \boldsymbol{x}H_2O$ in distilled water, making up the solution to 250.0 cm^3 in a volumetric flask and shaking the mixture.
- The student titrated this solution against 25.0 cm^3 portions of an acidified solution of 0.0200 mol dm^{-3} potassium manganate(VII), $KMnO_4(aq)$.

Method 2

- The student mixed a solution of $Mg(NO_2)_2 \bullet \boldsymbol{x}H_2O$ with an excess of aqueous sodium carbonate solution, $Na_2CO_3(aq)$.
- The student obtained a precipitate of magnesium carbonate, $MgCO_3(s)$, and determined the mass of this precipitate.

Method 3

- The student heated a known mass of $Mg(NO_2)_2 \bullet \boldsymbol{x}H_2O(s)$.
- The student determined the mass of the anhydrous residue formed.

Method 1 – Titration

The student filled the burette with the solution made from $Mg(NO_2)_2 \bullet \boldsymbol{x}H_2O$.

In each titration

- 25.0 cm^3 of 0.0200 mol dm^{-3} $KMnO_4(aq)$ was transferred to a conical flask using a pipette.
- An excess of dilute sulfuric acid was added to the conical flask and the mixture heated.
- $Mg(NO_2)_2(aq)$ was added from the burette until the end-point was reached.

The student's titration results are shown in the table (the rough titration results have **not** been included in the table).

Titration number	1	2	3
Final burette reading / cm^3	23.95	48.05	23.85
Initial burette reading / cm^3	0.80	24.50	0.65
Titre / cm^3			
Concordant titres (✓)			
Mean titre / cm^3			

(a) Complete the table.

(2)

(b) Deduce the colour change that the student would see at the end-point in this titration.

(1)

From ..... to
.....

(c) In the titration reaction, 2 mol MnO_4^- react with 5 mol NO_2^-.

Calculate the number of moles of NO_2^-, in the 250 cm^3 of solution prepared by the student and hence the value of $\boldsymbol{x}$ in $Mg(NO_2)_2 \bullet \boldsymbol{x}H_2O$.
Give your answer to the nearest whole number.

(5)

(d) The half-equations for the reaction in the titration are

$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

$NO_2^- + H_2O \rightarrow NO_3^- + 2H^+ + 2e^-$

Use these half-equations to derive the overall ionic equation for the reaction between manganate(VII) and nitrate(III) ions in acidic conditions.
State symbols are not required.

(2)

Method 2 – Precipitation

The student used the following procedure.

- Dissolve a known mass of $Mg(NO_2)_2 \cdot \boldsymbol{x}H_2O$ in distilled water.
- Add an excess of aqueous sodium carbonate solution, $Na_2CO_3(aq)$, to obtain a precipitate of magnesium carbonate, $MgCO_3(s)$.

$$Mg^{2}+(aq) + CO_3^{2-}(aq) \rightarrow MgCO_3(s)$$

- Weigh a piece of filter paper.
- Filter the mixture from the above reaction through the pre-weighed filter paper.
- Wash the precipitate of $MgCO_3(s)$ with distilled water.
- Dry the filter paper and precipitate in a desiccator.
- Reweigh the filter paper and the precipitate.
- Calculate the value of ***x*** from the results obtained.

The student found that the value of ***x*** calculated using **Method 2** was different from that obtained using **Method 1**. This difference occurred despite having used a pure sample of the hydrated salt and without making any errors in technique during the experiment.

The student found out from a data book that the compound magnesium carbonate is very slightly soluble in water.

(e) Explain how, if at all, the very slight solubility of magnesium carbonate in water would affect the value calculated for ***x***.

(2)

.....

.....

.....

.....

(f) The student planned to obtain any dissolved magnesium carbonate by evaporating the filtrate, and then weighing the residue.

Criticise this student's plan.

(2)

.....

.....

.....

.....

Method 3 – Thermal decomposition

NOTE: On heating, $Mg(NO_2)_2 \cdot \boldsymbol{x}H_2O(s)$ loses its water of crystallisation and **then** undergoes further decomposition to give magnesium oxide, MgO.

The student used the following procedure.

- Weigh an empty crucible.
- Add some $Mg(NO_2)_2 \cdot \boldsymbol{x}H_2O(s)$ and then reweigh the crucible plus contents.
- Heat the crucible plus contents and allow to cool.
- Weigh the crucible plus magnesium oxide residue.
- Use these data to calculate a value for ***x***.

The student's results are shown in the table.

Mass of crucible / g	18.02
Mass of crucible + $Mg(NO_2)_2 \cdot \boldsymbol{x}H_2O$ / g	18.84
Mass of crucible + MgO residue / g	18.27

(g) Identify how the student should ensure that the hydrated salt was fully decomposed.

(1)

.....

.....

(h) The student carried out an evaluation of the results obtained from **Method 3**

Identify **two** modifications to the method that would enable the student to lower the percentage uncertainty in the measurement of the mass of the solid residue.

(2)

.....

.....

.....

.....

(Total for question = 17 marks)

Q3. Which of the following physical methods of chemical analysis is used in modern breathalysers?

☒ **A** Infrared spectroscopy

☒ **B** Mass spectrometry

☒ **C** Nuclear magnetic resonance

☒ **D** Ultraviolet spectroscopy

(Total for Question = 1 mark)

Q4.

This question is about the chemistry of elements in the *d*-block of the Periodic Table.

Under certain conditions, dichromate(VI) ions, $Cr_2O_7^{2-}$, can oxidise manganese(II) ions, Mn^{2+}.

In this reaction, dichromate(VI) ions are reduced to chromium(III) ions, in acidic conditions, according to the half-equation

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

In an experiment it was found that 20.0 cm^3 of 0.100 mol dm^{-3} potassium dichromate(VI) was required to oxidise 30.0 cm^3 of 0.200 mol dm^{-3} manganese(II) sulfate solution.

Use these data to calculate the final oxidation state of the manganese.

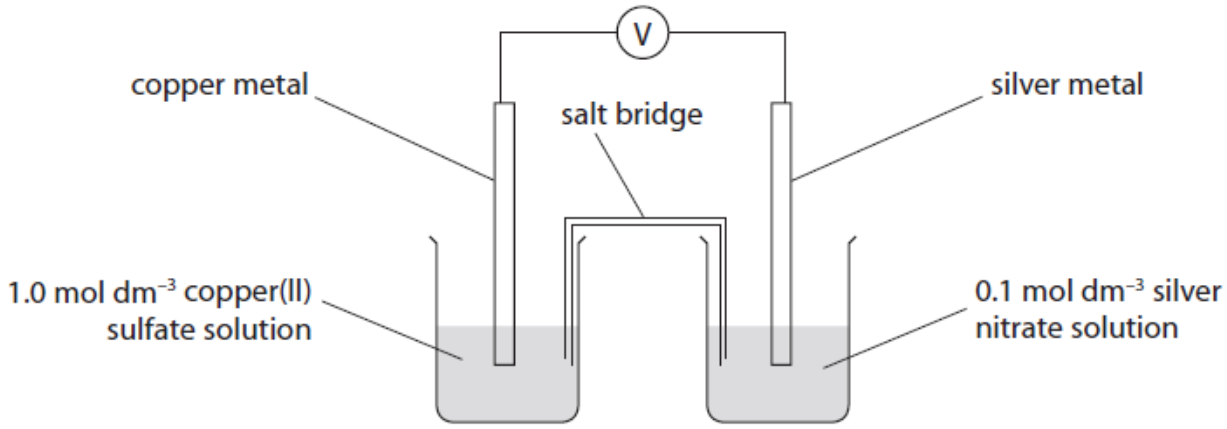
(5)

(Total for question = 5 marks)

Q5.

The diagram below shows a cell set up between a standard copper metal / copper(II) ion electrode and a silver metal / silver(I) ion electrode in which the silver ion concentration is 0.1

mol dm^{-3}.



At 298 K, the emf of this cell was +0.40 V. The electrode potential of the copper metal / copper(II) ion electrode is +0.34 V. What is the electrode potential of this silver metal / silver(I) ion electrode?

☒ **A** −0.74 V

☒ **B** −0.06 V

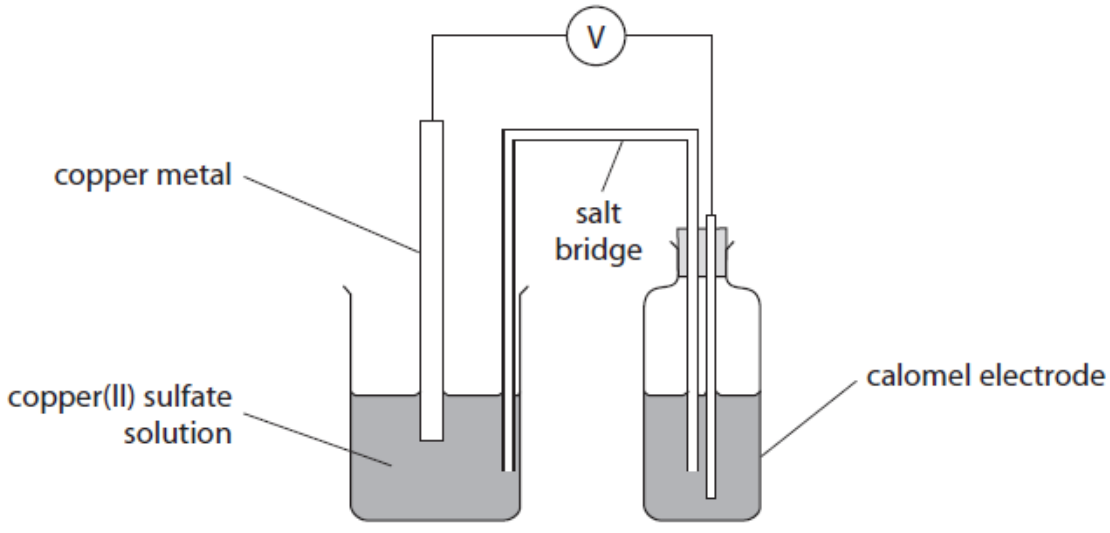
☒ **C** +0.06 V

☒ **D** +0.74 V

(Total for question = 1 mark)

Q6.

The diagram below shows a cell set up between a copper metal / copper(II) ion electrode and a reference electrode, known as a calomel electrode.



Under standard conditions, the emf of this cell was −0.07 V. The standard electrode potential of

the copper metal / copper(II) ion electrode is +0.34 V. Hence the standard electrode potential of the calomel electrode is

☒ **A** −0.41 V

☒ **B** −0.27 V

☒ **C** +0.27 V

☒ **D** +0.41 V

(Total for question = 1 mark)

Q7.

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)

.....

.....

.....

.....

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

(2)

.....

.....

.....

.....

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

(1)

.....

.....

.....

.....

(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)

.....

.....

.....

.....

(Total for question = 16 marks)

Q8.

Compound **A** is a green solid containing one cation and one anion. It dissolves in water to form a green solution.

(a) Give the formulae of **two** cations which could be responsible for the green colour in the solid.

(2)

.....

(b) A student adds dilute sodium hydroxide solution, drop by drop, to an aqueous solution of **A**. A green precipitate forms, which dissolves in excess sodium hydroxide solution to form a dark green solution.

(i) Give the formula of the cation in **A**.

(1)

.....

(ii) Give a formula for the green precipitate.

(1)

.....

(iii) Give a formula for the species present in the dark green solution.

(1)

.....

(c) The student adds a few drops of acidified potassium manganate(VII) solution to another sample of a solution of **A** in a test tube.

State the colour change that occurs.

(1)

.....

.....

(d) The student acidifies 2 cm^3 of a solution of **A** with dilute nitric acid in a test tube and then adds a few drops of aqueous silver nitrate. A white precipitate is formed.

(i) Give the formula of the anion in **A**.

(1)

.....

(ii) This test is usually followed by the addition of ammonia solution to test the solubility of the precipitate.

Explain why this procedure in not suitable to confirm the identity of the anion in **A**.

(2)

.....

.....

.....

.....

Q9.

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.

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

This redox reaction could be used in an electrochemical cell.

The cell half-equations are

$$2CO_2(g) + 2e^- \rightleftharpoons C_2O_4^{2-}(aq)$$

$$8H^+(aq) + MnO_4^-(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$$

Write a cell diagram for this cell using the conventional representation.

(2)

(Total for question = 2 marks)

Q10.

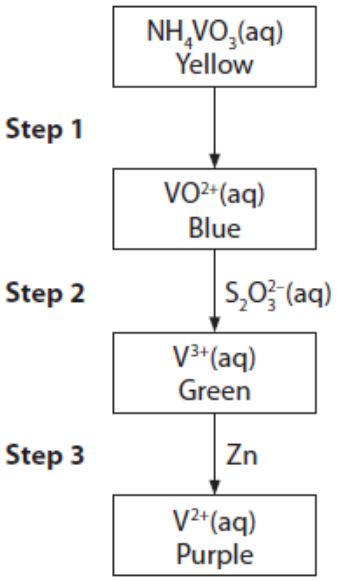
The element vanadium exists in a variety of oxidation numbers in its compounds.

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

(1)

.....

(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)

.....

.....

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

(1)

.....

.....

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

(2)

.....

.....

.....

.....

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

(2)

.....

.....

.....

.....

Q11.

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)

.....

.....

.....

.....

.....

.....

(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)

.....

.....

.....

.....

(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)

.....

.....

(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)

.....

.....

.....

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

(1)

.....

.....

(Total for question = 17 marks)

Q12.

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

Draw a labelled diagram of the apparatus that you would use to measure the standard emf of a cell with a zinc-zinc(II) electrode system and a chromium(III)-dichromate(VI) electrode system.

Include the **formulae** of all the compounds required and the concentrations of the solutions.

(7)

(Total for question = 7 marks)

Q13.

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.

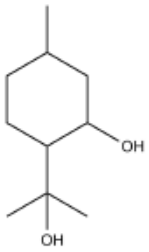
(3)

(Total for question = 6 marks)

Q14.

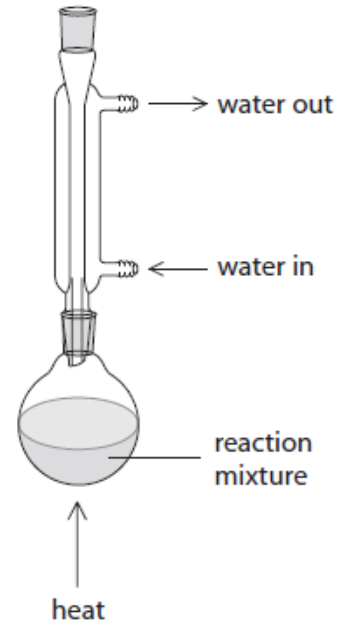
Alcohols contain the hydroxyl functional group. They can be oxidised using appropriate reagents and conditions.

Alcohol **Y** contains two hydroxyl groups.

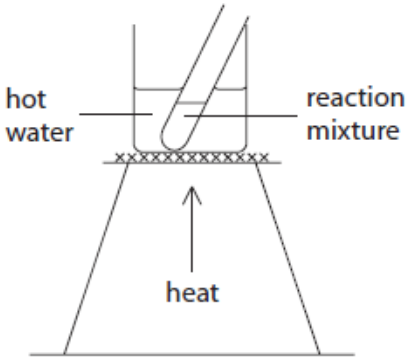


Alcohol **Y** can be oxidised by heating it with sodium dichromate(VI) solution and sulfuric acid.

*(a) A student suggested two possible techniques for heating the reaction mixture.



Technique A



Technique B

Evaluate all aspects of the two techniques and justify which one is more appropriate to just show that alcohol **Y** is able to be oxidised.

(6)

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

(b) Explain, using information from the Data Booklet, whether hydrochloric acid is a suitable alternative to sulfuric acid in the reaction mixture.

(2)

.....

.....

.....

.....

(c) Draw the **skeletal** formula of the organic product formed by the oxidation of alcohol **Y**.

(1)

(Total for question = 9 marks)

Q15.

The standard electrode potential, $E^{\ominus}$, of the $Ag^+(aq)|Ag(s)$ half-cell is +0.80 V.

The effect of changing the concentration of the ions on the value of the electrode potential, E, in this half-cell is calculated using the equation

$$E = E^{\ominus} + \frac{RT}{96500} \times \ln[Ag^+(aq)]$$

where T is the temperature in kelvin and R is the gas constant.

The electrode potential of a $Ag^+(aq)|Ag(s)$ half-cell was measured at 20 °C and found to be +0.72 V.

Calculate the concentration of silver ions, in mol dm^{-3}, in this half-cell.

(3)

(Total for question = 3 marks)

Q16.

A group of students analysed a hydrated salt with the formula $KH_3(C_2O_4)_{\mathbf{y}}.\ \mathbf{z}H_2O$ where **y** and **z** are whole numbers.

The students carried out experiments to determine the values of **y** and **z**.

(a) Experiment 1 – to determine the value of **y**

One student was provided with a 0.0235 mol dm^{-3} solution of the salt.

25.0 cm^3 portions of the salt solution were acidified with excess dilute sulfuric acid and heated to about 60 °C.

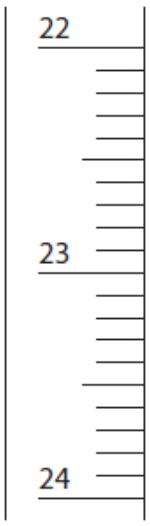
Each portion was titrated with 0.0203 mol dm^{-3} potassium manganate(VII).

The results of four titrations are shown in the table.

Titration number	1	2	3	4
Final burette reading / cm^3	23.85	47.20	24.05	48.10
Initial burette reading / cm^3	0.00	24.00	0.50	25.00
Titre / cm^3	23.85	23.20	23.55	23.10

(i) Complete the diagram to show the final burette reading in **Titration 1**.

(2)



(ii) Explain why this student should use a mean titre of 23.15 cm^3 and not 23.43 cm^3 in the calculation.

(2)

.....

.....

.....

.....

(iii) The uncertainty in each burette reading is ±0.05 cm^3.

Calculate the percentage uncertainty in the titre volume of potassium manganate(VII) solution used in **Titration 2**.

(1)

(iv) The equation for the reaction is

$$2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$$

Deduce, by calculation, the value of **y**, to the nearest whole number, in the formula $KH_3(C_2O_4)_{\mathbf{y}}.\ \mathbf{z}H_2O$.

Use the mean titre of 23.15 cm^3 and other data from **Experiment 1**.

You **must** show your working.

(4)

(b) **Experiment 2** – to determine the value of **z**

Another student wrote an account of the method for this experiment.

A crucible was weighed.

A sample of the hydrated salt was added to the crucible and it was reweighed.

The crucible and salt were heated to remove the water of crystallisation and then allowed to cool.

The crucible and contents were weighed again.

Results

Mass of crucible	*= 19.56 g*
Mass of crucible + $KH_3(C_2O_4)_{\mathbf{y}}.\mathbf{z}H_2O$	*= 22.97 g*
Mass of crucible + $KH_3(C_2O_4)_{\mathbf{y}}$	*= 22.52 g*

(i) Deduce, by calculation, the value of **z**, to the nearest whole number, in the formula $KH_3(C_2O_4)_{\mathbf{y}}.\ \mathbf{z}H_2O$.

You must use the data from **Experiment 2** and your value of **y** in (a)(iv).

You **must** show your working.

(3)

(ii) A third student carried out Experiment 2 and calculated a value of **z** that was lower than expected.

This student evaluated the experiment and gave two suggestions for **z** being lower.

Suggestion **1**
"Some of the crystals jumped out of the crucible while it was being heated."

Suggestion **2**
"It was difficult to tell when all the water of crystallisation had been lost."

Evaluate these two suggestions to decide whether they could account for the lower value of **z** obtained from the experimental results.

Include an explanation of the effect each suggestion would have on the calculated value of **z** and how the method could be improved to prevent these errors.

(5)

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

(Total for question = 17 marks)

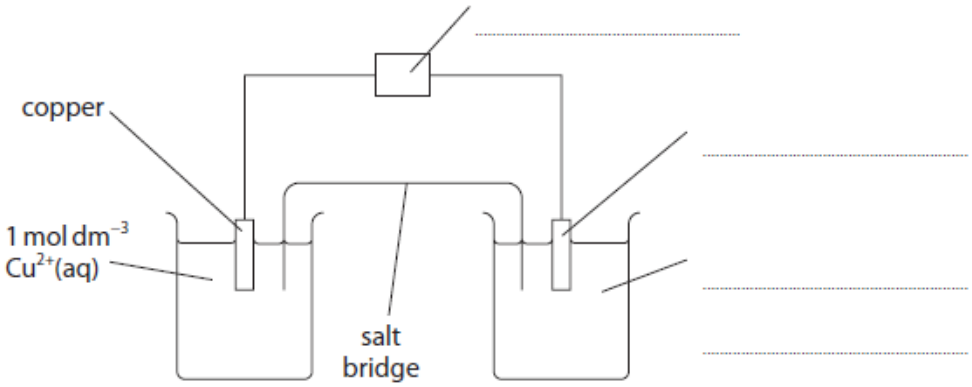
Q17.

An electrochemical cell is made from the electrode systems represented by these half-equations.

$$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$$

$$Mn^{3+}(aq) + e^- \rightleftharpoons Mn^{2+}(aq)$$

The $E^{\ominus}_{cell}$ value is measured using the apparatus shown.



(a) Complete the diagram by adding labels on the dotted lines provided.

(3)

(b) A salt bridge is used to connect the two half-cells.

(i) State what chemical is contained in the salt bridge.

(1)

.....

.....

(ii) Give a possible reason why the salt bridge cannot be replaced by an unreactive metal wire.

(1)

.....

.....

.....

(c) In this cell, the copper is oxidised and $E^{\ominus}_{cell}$ = +1.15 V.

$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$ $\quad E^{\ominus} = +0.34\,V$

$Mn^{3+}(aq) + e^- \rightleftharpoons Mn^{2+}(aq)$

(i) Write the overall ionic equation for the reaction taking place.

State symbols are not required.

(1)

(ii) Calculate the value of the standard electrode potential for the $Mn^{3+}(aq)$ I $Mn^{2+}(aq)$ half-cell.

(1)

(Total for question = 7 marks)

Q18. In aqueous solution, manganate(VI) ions disproportionate into manganate(VII) ions and manganese(IV) oxide when carbon dioxide is bubbled through the solution. The ionic equation for the reaction is

$$3MnO_4^{2-}(aq) + 4H^+(aq) \rightarrow 2MnO_4^-(aq) + MnO_2(s) + 2H_2O(l)$$

The role of the carbon dioxide is to

☒ **A** lower the pH of the solution.

☒ **B** raise the pH of the solution.

☒ **C** oxidize the manganate(VI) ions.

☒ **D** reduce the manganate(VI) ions.

(Total for Question = 1 mark)

Q19. 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.

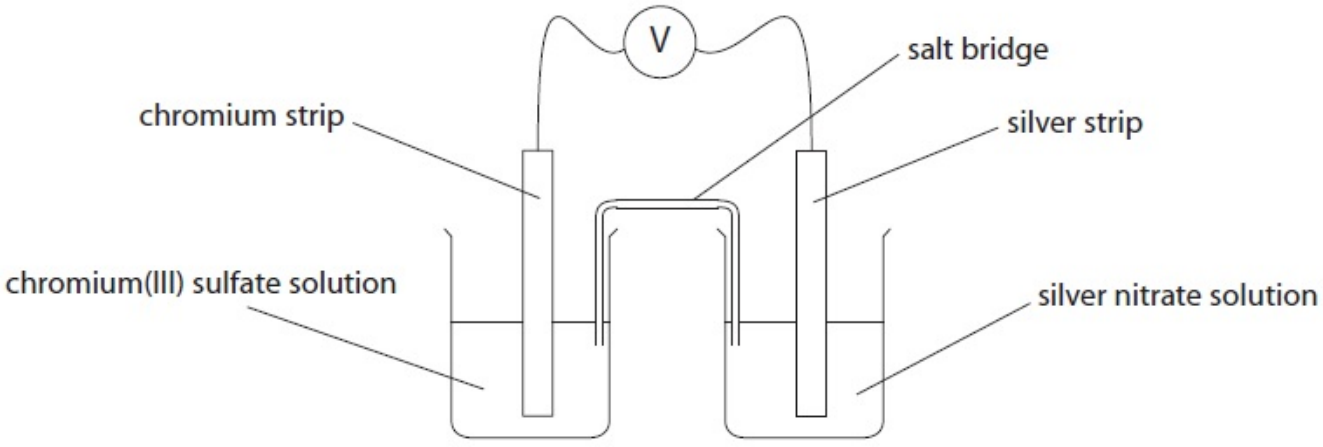
☒ **C** 1 mol of manganate(VII) ions reacts with 2.5 mol of chlorine.

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

(Total for Question = 2 marks)

Q20.

The diagram below shows an electrochemical cell.



(a) The salt bridge in this cell is **best** prepared by soaking filter paper in a

⊖ **A** 1 mol dm^{-3} solution of potassium bromide.

- **B** 1 mol dm^{-3} solution of potassium chloride.
- **C** saturated solution of potassium nitrate.
- **D** saturated solution of potassium iodide.

(1)

(b) The relevant standard electrode potentials for this cell are shown below.

$Cr^{3+}(aq) + 3e^- \rightleftharpoons Cr(s) \quad E^\ominus = -0.74\ V$

$Ag+^-(aq) + e^- \rightleftharpoons Ag(s) \quad E^\ominus = +0.80\ V$

(1)

The emf of the cell shown in the diagram is

- **A** +1.54 V
- **B** −1.54 V
- **C** +0.06 V
- **D** −0.06 V

(c) The mass of 1 mol of chromium(III) sulfate ($Cr_2(SO_4)_3$) is 392 g. Hence, for the cell in the diagram to measure the **standard** electrode potential, 1 dm^3 of the chromium(III) sulfate solution used must contain

(1)

- **A** 104 g of chromium(III) sulfate.
- **B** 196 g of chromium(III) sulfate.
- **C** 392 g of chromium(III) sulfate.
- **D** 784 g of chromium(III) sulfate.

(Total for question = 3 marks)

Q21.

This question is about chlorine.

The standard electrode potential for the chlorine / chloride ion half-cell is

$$\tfrac{1}{2}Cl_2(aq) + e^- \rightleftharpoons Cl^-(aq) \quad E^\ominus_{cell} = +1.36\ V$$

(i) Identify an oxidising agent from the Data Booklet that will convert chloride ions into chlorine under standard conditions.

(1)

.....

(ii) Calculate the value of $E^{\ominus}_{cell}$ for the reaction in (i).

(1)

(Total for question = 2 marks)

Q22.

For the reaction

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

$E^{\ominus}_{cell}$ is positive. From this it can be deduced that, for this reaction,

ΔS_{total} **A** ΔS_{total} and ln$K$ are positive.

ΔS_{total} **B** ΔS_{total} and ln$K$ are negative.

ΔS_{total} **C** ΔS_{total} is positive and ln$K$ is negative.

ΔS_{total} **D** ΔS_{total} is negative and ln$K$ is positive.

(Total for question = 1 mark)

Q23. The equation for the reaction of iron and nickel(II) ions in aqueous solution is

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

Under standard conditions the value of the equilibrium constant, K_c, for this reaction is greater than 1. Hence, for this reaction,

☒ **A** $\Delta S^{\ominus}_{total}$ and $E^{\ominus}_{reaction}$ are both positive.

☒ **B** $\Delta S^{\ominus}_{total}$ is positive and $E^{\ominus}_{reaction}$ is negative.

☒ **C** $\Delta S^{\ominus}_{total}$ is negative and $E^{\ominus}_{reaction}$ is positive.

☒ **D** $\Delta S^{\ominus}_{total}$ and $E^{\ominus}_{reaction}$ are both negative.

(Total for Question = 1 mark)

Q24.

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

Excess zinc powder is added to an acidified solution of the compound NH_4VO_3. Using the data in the table, explain the sequence of reactions that takes place.

In your answer, include a description of what you would **see**, and the relevant ionic equations with their calculated $E^{\ominus}_{cell}$ values. State symbols are not required.

(7)

Electrode system	$E^{\ominus}$ / V
$V^{2+}(aq) + 2e^- \rightleftharpoons V(s)$	−1.18
$V^{3+}(aq) + e^- \rightleftharpoons V^{2+}(aq)$	−0.26
$VO^{2+}(aq) + 2H^+(aq) + e^- \rightleftharpoons V^{3+}(aq) + H_2O(l)$	+0.34
$VO_3^-(aq) + 4H^+(aq) + e^- \rightleftharpoons VO^{2+}(aq) + 2H_2O(l)$	+1.00
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	−0.76

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

(Total for question = 7 marks)

Q25.

The table shows the standard electrode (redox) potentials, $E^{\ominus}$, for some half-cell reactions.

Redox system	Half-cell reaction	$E^{\ominus}$ / V
1	$V^{2+}(aq) + 2e^- \rightleftharpoons V(s)$	−1.20
2	$V^{3+}(aq) + e^- \rightleftharpoons V^{2+}(aq)$	−0.26
3	$VO^{2+}(aq) + 2H^+(aq) + e^- \rightleftharpoons V^{3+}(aq) + H_2O(l)$	+0.34
4	$VO_2^+(aq) + 2H^+(aq) + e^- \rightleftharpoons VO^{2+}(aq) + H_2O(l)$	+1.00
5	$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons SO_2(g) + 2H_2O(l)$	+0.17

(i) Explain, using information from the table, the colour changes that take place when SO_2 gas is bubbled slowly through an acidified solution containing VO_2^+ ions.

Equations are not required.

(3)

.....

.....

.....

.....

.....

.....

(ii) Explain, using information in the table, whether the disproportionation of $V^{2+}(aq)$ into $V^{3+}(aq)$ and V(s) is feasible under standard conditions.

(2)

.....

.....

.....

.....

.....

.....

Q26. Hydrogen combines rapidly with oxygen in the presence of a platinum catalyst:

$$H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(l)$$

The reaction is highly exothermic.

(a) Use the thermochemical data from the data booklet to obtain the enthalpy change for this reaction under standard conditions.

(1)

..... kJ mol^{-1}

(b) The same reaction occurs, also with a platinum catalyst, in a fuel cell.

(i) Write the two ionic half equations which occur in an **alkaline** fuel cell. Include state symbols.

(3)

Equation 1

Equation 2

(ii) The alkali in an alkaline fuel cell serves the same purpose as the acid in an acid fuel cell. State this purpose.

(1)

.....

.....

*(iii) Platinum catalyses both the direct combination of hydrogen with oxygen and the reactions in the fuel cell. By considering the way in which the catalyst lowers the activation energy, suggest two **similarities** in these processes.

(2)

Similarity 1

.....

.....

.....

Similarity 2

.....

.....

.....

(c) The use of hydrogen as a fuel, both in fuel cells and in direct combustion reactions, is seen as an important potential alternative to fossil fuels.

(i) State what is considered to be the main advantage of hydrogen compared with fossil fuels, bearing in mind that most hydrogen is obtained from fossil fuels.

(1)

.....

.....

*(ii) Explain the main advantage of using a fuel cell over direct combustion of hydrogen.

(2)

.....

.....

.....

.....

(iii) State a disadvantage of using a hydrogen fuel cell compared with direct combustion of hydrogen.

(1)

.....

.....

(iv) Suggest **two** advantages of using an ethanol fuel cell rather than a hydrogen fuel cell.

(2)

.....

.....

.....

.....

(Total for Question = 13 marks)

Q27.

In a hydrogen-oxygen fuel cell, hydrogen is

☒ **A** oxidized at the anode.

☒ **B** oxidized at the cathode.

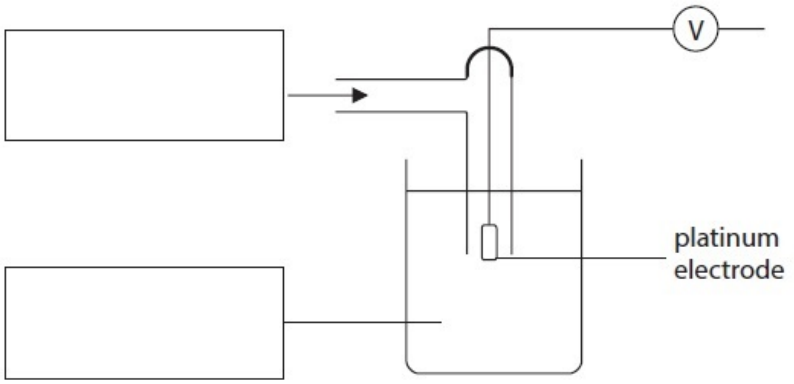
☒ **C** reduced at the anode.

☒ **D** reduced at the cathode.

(Total for question = 1 mark)

Q28. Hydrogen-oxygen fuel cells can operate in acidic or alkaline conditions. One such commercial cell uses porous platinum electrodes in contact with concentrated aqueous potassium hydroxide solution, KOH(aq).

(a) Use relevant standard electrode potential values, on pages 15 and 17 of the Data Booklet, to complete the table below in which two *E*



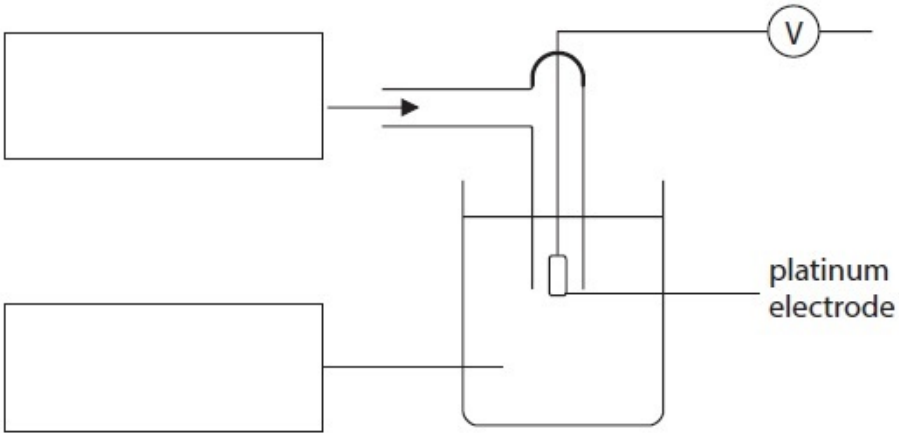
values are missing.

(2)

Half-equation	$E^{\ominus}$ / V
$2H_2O(l) + 2e^- \rightleftharpoons 2OH^-(aq) + H_2(g)$	−0.83
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$	
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	

(b) (i) Fill in the boxes to identify, by name or formula, the substances used in the **standard** hydrogen electrode.

(2)



(ii) State **three** conditions that are necessary for a standard hydrogen electrode.

(2)

1. .....

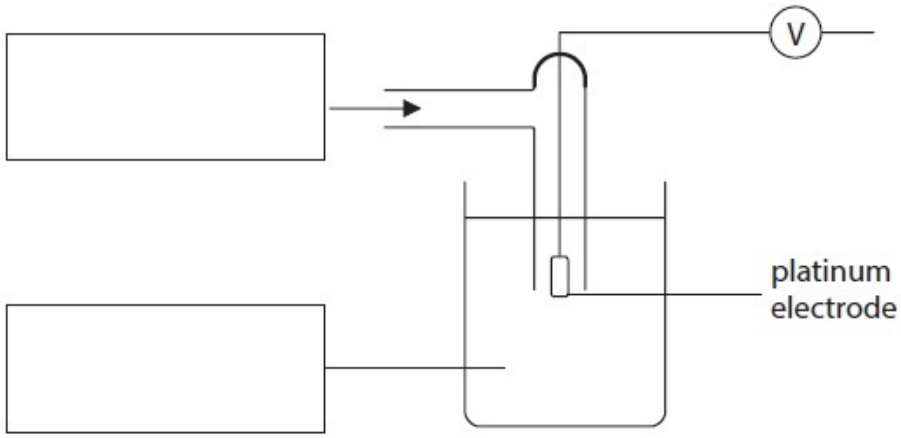
2. .....

3. .....

(c) Write appropriate half-equations and use them to derive an overall equation for the reaction which occurs when an **alkaline** hydrogen-oxygen fuel cell operates.

(2)

(d) Use the *E*



values from the table in part (a) to calculate the $E^{\ominus}_{cell}$ for a hydrogenoxygen fuel cell operating in alkaline conditions.

(1)

(e) Suggest why the $E^{\ominus}_{cell}$ cell for a hydrogen-oxygen fuel cell, operating in **acidic** conditions, is identical to that of an alkaline fuel cell.

(1)

.....

.....

(f) Give **one** reason (other than cost implications) why the platinum electrodes are made by coating porous material with platinum rather than by using platinum rods.

(1)

.....

.....

(g) Suggest **one** disadvantage of using a hydrogen-oxygen fuel cell compared with a rechargeable battery when providing electrical energy for a motor vehicle.

(1)

.....

.....

(Total for Question = 12 marks)

Q29.

This question is about the chemistry of elements in the *d*-block of the Periodic Table.

A student constructed an electrochemical cell as follows:

- a half-cell was made from a strip of chromium metal and a solution of aqueous chromium(III) sulfate
- a second half-cell was made from a piece of metal, **X**, and a solution of its sulfate, **X** SO_4(aq)
- the two half-cells were connected and a current allowed to pass for some time.

Results

- the chromium electrode increased in mass by 1.456 g
- the electrode made of metal **X** decreased in mass by 1.021 g.

Use these data to determine the identity of the metal, **X**.

(4)

(Total for question = 4 marks)

Q30.

This question is about redox reactions.

Identify the species that is the strongest reducing agent from the list of standard electrode potentials in the Data Booklet.

(1)

.....

(Total for question = 1 mark)

Q31.

If the concentration of chlorate(I) ions in a swimming pool is too low, bacteria will flourish. If it is too high, swimmers will be harmed. In the United Kingdom, it is recommended that the concentration of chlorine be in the range 1–2 mg dm^{-3}.
The total amount of chlorine, Cl_2, present can be measured by titration.

In such a determination, excess sulfuric acid and potassium iodide were added to a 1.00 dm^3 sample of swimming pool water and the resulting solution required 9.65 cm^3 of 0.00550 mol dm^{-3} sodium thiosulfate for complete reaction.

The equations for the reactions are:

$$Cl_2 + 2I^- \rightarrow I_2 + 2Cl^-$$

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

(i) Show that the concentration of chlorine, Cl_2, in this sample meets the UK recommended requirements.
All steps in your calculation must be shown.

(3)

(ii) Analysing a single sample of swimming pool water in this way is likely to give unreliable results because it is not possible to repeat the titration.
Suggest another way in which this method is unreliable.

(1)

.....

.....

Q32.

Sulfur dioxide is used as a preservative in wine. It kills bacteria and helps to prevent premature oxidation of the alcohol. However, if the concentration of the sulfur dioxide is too high, the taste and smell of the wine are adversely affected. The European Wine Regulations fix the maximum concentration for sweet white wine at 400 mg dm^{-3}.

The concentration of sulfur dioxide in a sample of wine can be determined by the Ripper method, which makes use of a technique known as back-titration.

Sulfur dioxide and iodine react in aqueous solution according to the equation:

$$SO_2(aq) + I_2(aq) + 2H_2O(l) \rightarrow H_2SO_4(aq) + 2HI(aq)$$

A 10.0 cm^3 sample of sweet white wine is treated with 10.0 cm^3 of 0.00500 mol dm^{-3} (an excess) of iodine solution.

After the reaction with sulfur dioxide is complete, the remaining iodine is titrated with 0.00100 mol dm^{-3} sodium thiosulfate, using a suitable indicator.

38.70 cm^3 of the sodium thiosulfate solution is required.

(a) Name a suitable indicator for this titration and state the colour change observed at the end-point.

(2)

.....

.....

(b) Complete the ionic equation for the reaction between sodium thiosulfate and iodine. State symbols are not required.

(1)

$$2S_2O_3^{2-} + I_2 \rightarrow \text{.....} + \text{.....}$$

(c) Use the data provided to deduce whether this sweet white wine may be sold within the European Union.

(6)

.....

(Total for question = 9 marks)

Q33. The percentage by mass of tin in a piece of rock containing tin(IV) oxide, SnO_2, was determined as described in the procedure below.

Step 1 A sample of rock, with mass 10.25 g, was crushed and dissolved in sulfuric acid.

Step 2 The solution was treated with a reducing agent to convert the Sn^{4+} to Sn^{2+} ions.

Step 3 50 cm^3 of aqueous iodine solution with concentration 0.250 mol dm^{-3} was added to the solution of Sn^{2+} ions. The following reaction occurred:

$$Sn^{2+}(aq) + I_2(aq) \rightarrow Sn^{4+}(aq) + 2I^-(aq)$$

Step 4 The **excess** iodine was titrated with sodium thiosulfate solution with concentration 0.100 mol dm^{-3}. The volume of sodium thiosulfate solution required was 11.60 cm^3.

(a) Thiosulfate ions react with iodine as shown below.

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

(i) Calculate the number of moles of sodium thiosulfate which were used in **Step 4**.

(1)

(ii) Calculate the number of moles of iodine which reacted with this amount of sodium thiosulfate.

(1)

(iii) Calculate the number of moles of iodine added to the solution of Sn^{2+} ions in **Step 3**.

(1)

(iv) Use your results from (ii) and (iii) to calculate the number of moles of iodine which reacted with the Sn^{2+} ions from the rock.

(1)

(v) Hence calculate the percentage by mass of tin in the rock.

(2)

(b) (i) What change could be made in **Step 4** to improve the reliability of the result?

(1)

.....

.....

.....

(ii) The error each time the burette was read was ± 0.05 cm^3. Calculate the percentage error in

the titre value of 11.60 cm^3.

(1)

(iii) How could the percentage error in the titre value be reduced without using a different burette?

(1)

.....

.....

(c) The titration can be carried out with or without an indicator. What colour **change** would be seen at the end-point if an indicator was **not** used? The tin ions are colourless.

(1)

.....

(Total for Question = 10 marks)

Q34.

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)

.....

.....

.....

.....

(Total for question = 3 marks)

Q35.

Yellow gold is used to make jewellery. It is an alloy of copper, gold and silver. The purity of gold is measured in carats. The higher the carat, the higher the percentage of gold in the alloy. Pure gold is 24 carat.

A sample of yellow gold is analysed using the steps below.

Step **1** Excess concentrated nitric acid is reacted with 1.250 g of the alloy. The gold does **not** react but the copper and silver do react. The half-equations are

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

$$Ag(s) \rightarrow Ag^+(aq) + e^-$$

$$2HNO_3(aq) + e^- \rightarrow NO_3^-(aq) + NO_2(g) + H_2O(l)$$

Step **2** The mixture is diluted with distilled water and the gold is filtered off.

Step **3** Excess hydrochloric acid is added to the filtrate. It reacts with the silver ions to form a precipitate of silver chloride.

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

Step **4** The silver chloride precipitate is filtered off, washed, dried and weighed. The mass of silver chloride formed is 0.706 g.

Step **5** Excess potassium iodide is added to the remaining solution. A precipitate of copper(I) iodide and a solution of iodine forms.

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

Step **6** The resulting mixture is titrated with 0.100 mol dm^{-3} sodium thiosulfate solution.

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

The titre is 39.40 cm^3.

The table shows the percentage by mass of gold in four different carats of yellow gold.

Carat	Percentage by mass of gold
9	37.5
10	41.7
14	58.3
18	75.0

Determine, using the experimental data, the carat of the sample of yellow gold that was analysed.

(6)

(Total for question = 6 marks)

Q36.

Complexes involve ligands forming dative covalent bonds with a central cation; the number of dative covalent bonds is the coordination number of the cation in the complex. Complexes may be positive, negative or neutral. The ligands may be neutral or negatively charged, but are never positively charged. The ligands must have a lone pair of electrons. In some complexes the central atom is neutral, as is the case with nickel carbonyl which is used in the Mond process for the purification of nickel:

$$Ni(s) + 4CO(g) \rightleftharpoons Ni(CO)_4(g)$$

The formation of complexes is typical of transition metals but other elements also form complexes. Zinc, which is a d block element but not a transition metal, and aluminium, which is a Group 3 element, both form complexes. Transition metal complexes are usually coloured, but if the interaction between the ligand and the central ion is very strong, then the resulting complex may well be colourless. This is the case with the complex $[FeF_6]^{3-}$.

The ligand in a complex affects its redox properties and this is apparent from considering the relevant standard electrode potentials. For example, with copper(II) species:

$$[Cu(H_2O)_6]^{2+}(aq) + 2e^- \rightleftharpoons Cu(s) + 6H_2O(l) \quad E^{\ominus} = +0.34\,V$$

$$[Cu(H_2O)_2(NH_3)_4]^{2+}(aq) + 2e^- \rightleftharpoons Cu(s) + 2H_2O(l) + 4NH_3(aq) \quad E^{\ominus} = -0.05\,V$$

Complexes are industrially important, for example in the purification of nickel mentioned above and in the extraction of gold which involves the complex $[Au(CN)_2]^-$.

Complex formation is used in estimating the concentrations of metal ions in solution. Examples of this are the use of EDTA in titration and dimethylglyoxime in gravimetric analysis. Complexes are also used in qualitative analysis, for example, in Tollens' reagent. This reagent is usually prepared by the addition of sodium hydroxide solution to an aqueous solution of silver nitrate followed by aqueous ammonia. The resulting solution contains diamminesilver(I) ions.

(a) The hydrated copper(II) ion is classed as a complex but the hydrated sodium ion is not. By considering the interaction between the water molecules and the sodium ion, explain why the hydrated sodium ion is **not** considered to be a complex.

(2)

.....

.....

.....

.....

(b) Suggest how the aluminium ion is able to form dative covalent bonds in its complexes such

as $[AlF_6]^{3-}$.

(1)

.....

.....

.....

*(c) (i) Transition metal complexes are usually coloured. Explain how the colours occur.

(4)

.....

.....

.....

.....

.....

.....

.....

.....

(ii) Why are zinc complexes such as $[Zn(NH_3)_4]^{2+}$ colourless?

(1)

.....

.....

(iii) Suggest why the strong interactions between the F^- ligands and Fe^{3+} ion result in the complex $[FeF_6]^{3-}$ being colourless.

(2)

.....

.....

.....

.....

(d) Use the standard electrode potentials given in the passage to predict whether it is feasible for thiosulfate ions, $S_2O_3^{2-}$, to reduce copper(II) ions to copper. Consider the case when the copper(II) ions are dissolved in water, and when they are dissolved in aqueous ammonia. Give chemical equations to support your answer.

(3)

The relevant redox reaction for thiosulfate ions is

$$S_4O_6^{2-}(aq) + 2e^- \rightleftharpoons 2S_2O_3^{2-}(aq) \quad E^{\ominus} = +0.09\,V$$

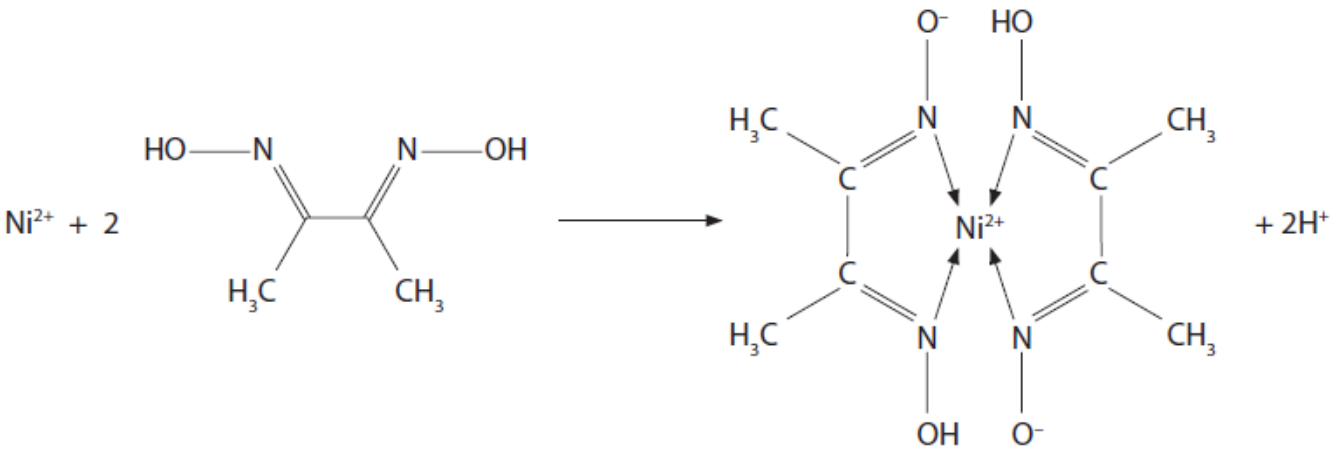
.....

.....

.....

(e) The proportion of nickel in nisil, an alloy of nickel and silicon used in thermocouples, may be determined by the following gravimetric analysis.

Finely ground nisil is dissolved in concentrated hydrochloric acid and, after neutralization, excess dimethylglyoxime dissolved in ethanol is added to the solution formed. The resulting red precipitate is filtered, washed, dried and weighed. The equation for the formation of the precipitate is



(i) State the coordination number of the nickel in the complex ion.

(1)

.....

(ii) In an experiment, 1.02 g of nisil formed 4.82 g of the complex. Calculate the percentage by mass of nickel in the alloy. The molar mass of the complex is 288.7 g mol^{-1}.

(2)

(Total for question = 16 marks)

Q37.

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)

.....

.....

.....

.....

(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** $Fe^{2+}(aq)$

Q38.

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** $Mn^{2+}(aq)$

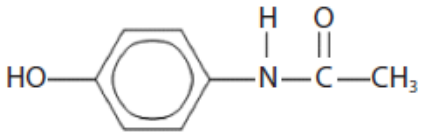
(ii) Write the ionic equation and calculate the $E^{\ominus}_{cell}$ for the reaction between MnO_4^- ions and Br^- ions in acidic solution.
State symbols are not required.

(3)

(Total for question = 4 marks)

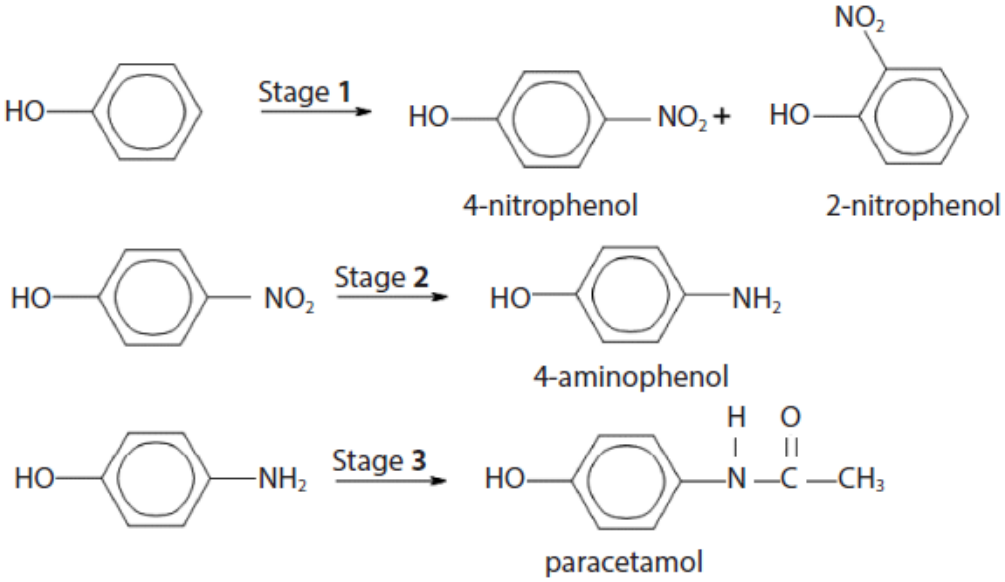
Q39.

This question is about the preparation and analysis of paracetamol.



paracetamol

Paracetamol may be prepared from phenol in three stages.

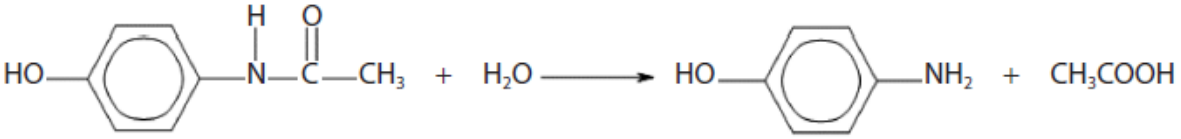


The percentage by mass of paracetamol in three brands of paracetamol tablets are shown in the table.

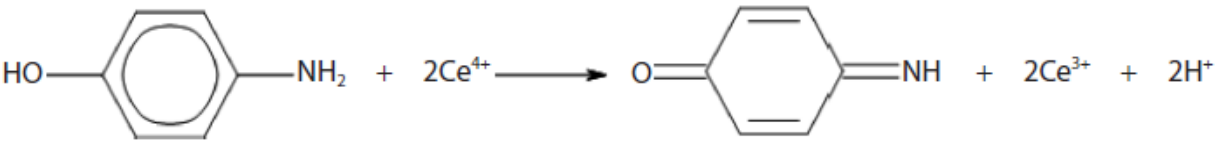
Brand of tablet	Percentage by mass of paracetamol
P	92.1
Q	93.6
R	99.7

The amount of paracetamol in a tablet can be determined using a titration with cerium(IV) ions.

The tablets are crushed and then hydrolysed in acid to form 4-aminophenol.



4-aminophenol is oxidised by cerium(IV) ions.



A tablet from one of the three brands of paracetamol was analysed following the outline procedure.

- one of the tablets was crushed and 0.500 g of the powder was added to dilute sulfuric acid
- the mixture was heated under reflux until the hydrolysis was complete
- the solution was made up to 100.0 cm^3 in a volumetric flask
- 25.0 cm^3 portions of the solution were titrated against 0.100 mol dm^{-3} Ce^{4+} using ferroin as indicator

Result

The mean titre was 16.50 cm^3.

(i) Give the molecular formula of paracetamol.

(1)

.....

(ii) Determine, by calculation, which brand of tablet was analysed.

(5)

(Total for question = 6 marks)

Q40.

A solution of potassium manganate(VII) was used to determine the concentration of iron(II) ions in solution by titration in the presence of excess dilute sulfuric acid.

(a) With the potassium manganate(VII) in the burette, the end-point of the reaction is when the solution in the conical flask turns

(1)

☒ **A** colourless.

☒ **B** pink.

☒ **C** green.

☒ **D** orange.

(b) If insufficient acid is added, the titre value is

(1)

☒ **A** low and a brown precipitate forms.

☒ **B** low and a green precipitate forms.

☒ **C** high and a brown precipitate forms.

☒ **D** high and a green precipitate forms.

(Total for question = 2 marks)

Q41.

The standard electrode potentials of two half reactions are shown below.

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

$$Co^{3+} + e^- \rightleftharpoons Co^{2+} \qquad E^{\ominus} = +1.82\,V$$

Which of the following processes is thermodynamically favourable? The reaction of

☒ **A** Co^{2+} with Cl_2 to form Cl^-

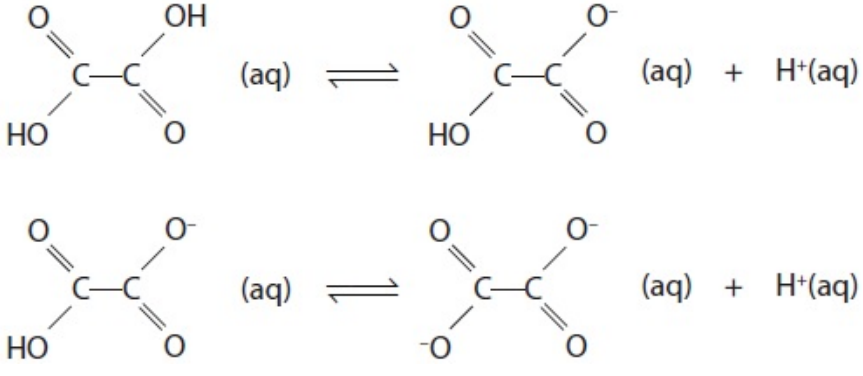
☒ **B** Co^{2+} with Cl^- to form Cl_2

☒ **C** Co^{3+} with Cl_2 to form Cl^-

☒ **D** Co^{3+} with Cl^- to form Cl_2

(Total for question = 1 mark)

Q42. Ethanedioic acid (also known as oxalic acid) is a dicarboxylic acid. Like sulfuric acid, it has two replaceable hydrogen atoms so it dissociates in two stages which are shown below. Acids like these are called 'dibasic'.



As is always the case with dibasic acids, the first dissociation occurs to a far greater extent than

the second. Ethanedioic acid is a carboxylic acid but, because there are two carboxyl groups, it is a much stronger acid than ethanoic acid.

Ethanedioic acid is toxic, mainly due to the precipitation of calcium ethanedioate in the kidneys. As it occurs in many plants, the toxicity of ethanedioic acid has caused a number of deaths notably during the First World War when people in England used rhubarb leaves, which contain significant concentrations of the compound, as a vegetable substitute. As it is a reducing agent, the concentration of ethanedioic acid is usually determined by titration with a solution of potassium manganate(VII) of known concentration.

The ethanedioate ion, which is formed when ethanedioic acid dissociates fully, is a bidentate ligand. The ethanedioate ion together with 1,2-diaminocyclohexane, another bidentate ligand, forms a complex with platinum(II) called oxaliplatin which has the same geometry as cisplatin. Oxaliplatin is an anti-cancer drug, which has been used to treat cancers that have become resistant to cisplatin.

Ethanedioic acid has numerous uses, from cleaning and rust removal to the purification of metals. Although it can be prepared from ethene in a two stage process, it is usually manufactured by the oxidation of carbohydrates or glucose using nitric acid, or air, in the presence of vanadium(V) oxide.

(a) (i) Suggest why the presence of the second carboxylic acid group increases the extent of the **first** dissociation of ethanedioic acid compared with the dissociation of ethanoic acid.

(2)

.....

.....

.....

.....

(ii) Explain, by referring to the equilibria in the passage, why the second dissociation of a dibasic acid occurs to a lesser extent than the first.

(1)

.....

.....

.....

.....

(b) In an experiment to determine the proportion of ethanedioic acid in rhubarb leaves, 250 g of leaves was crushed and boiled in a solution of sodium carbonate. The resulting mixture was filtered and the filtrate transferred, along with washings from the mixture, to a 250 cm^3 volumetric flask. The solution was acidified with sulfuric acid, made up to the mark and mixed thoroughly. 25.0 cm^3 portions of the solution were pipetted into a conical flask, about 25 cm^3 of sulfuric acid added and the mixture titrated against a solution of potassium manganate(VII) of concentration 0.0200 mol dm^{-3}. The mean titre was 28.55 cm^3. The relevant ionic half equations for the reaction are

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

$$C_2O_4^{2-}(aq) \rightarrow 2CO_2(g) + 2e^-$$

(i) State the colour change that occurs at the end-point of the titration.

(1)

.....

*(ii) Calculate the percentage by mass of ethanedioic acid in rhubarb leaves.

(5)

(c) (i) Explain the term 'bidentate'.

(1)

.....

.....

(ii) Draw a diagram of the oxaliplatin(II) complex, showing clearly the geometry and charge of the complex and how the ligands are attached to the central ion.

(3)

(d) (i) Suggest a two step **laboratory** synthesis of ethanedioic acid from ethene. Indicate the intermediate compound formed and the reagents required.

(3)

(ii) Suggest why it is preferable from an environmental perspective that ethanedioic acid is manufactured from carbohydrates and glucose rather than from ethene.

(1)

.....

.....

(iii) A sample of ethanedioic acid, manufactured from glucose, was contaminated with a small amount of propanoic acid.

High resolution nmr may be used to distinguish propanoic acid from ethanedioic acid.

Describe the high resolution nmr spectra of both propanoic acid and ethanedioic acid. Do **not** include chemical shift data in your answer.

(4)

.....

.....

.....

.....

.....

.....

.....

.....

(Total for Question = 21 marks)

Q43.

Which of the statements about a standard hydrogen electrode, for which $E^{\ominus} = 0$ V, is correct?

☒ **A** A suitable solution for use in the electrode is hydrochloric acid with a concentration of 0.1 mol dm^{-3}.

☒ **B** The pressure of the hydrogen has no effect on the value of $E^{\ominus}$.

☒ **C** The metal used in the electrode is platinum.

☒ **D** The temperature is 273 K.

(Total for question = 1 mark)

Q44.

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)

.....

.....

.....

.....

.....

.....

(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)

.....

(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)

.....

.....

.....

.....

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

(1)

.....

.....

.....

.....

Q45.

(a) When iron(II) sulfate is dissolved in water, the iron(II) ions are slowly oxidized to iron(III) ions by oxygen dissolved in the water.

(i) Write the two ionic half-equations for this redox reaction in **acid** conditions.
State symbols are not required.

(1)

(ii) Hence write the overall ionic equation for the reaction. State symbols are not required.

(1)

(b) 6.90 g of iron(II) sulfate crystals ($FeSO_4 \bullet 7H_2O$) was dissolved in distilled water and the solution made up to 250 cm^3 in a volumetric flask. After 24 hours, 25.0 cm^3 portions of this solution were pipetted into a conical flask and titrated against acidified potassium manganate(VII) solution with a concentration of 0.0195 mol dm^{-3}. The mean titre was 24.90 cm^3.

(i) Write the ionic equation showing that 1 mol of manganate(VII) ions oxidizes 5 mol of iron(II) ions in acid conditions. State symbols are not required.

(1)

(ii) What is the colour of the solution in the conical flask at the end-point of the titration?

(1)

.....

* (iii) Calculate the amount of Fe^{2+} ions in 250 cm^3 of the solution after it had been left to stand for 24 hours. Hence calculate the percentage of the iron(II) ions that had been oxidized between the preparation of the solution and the titration. The molar mass of $FeSO_4 \bullet 7H_2O$ is 277.9 g mol^{-1}.

(5)

(iv) Suggest, with an explanation, the appropriate number of significant figures to give for the answer to (b)(iii).

(1)

.....

.....

.....

(c) The most stable oxidation states in iron compounds are +2 and +3, but others do exist, for example in the ferrate(VI) ion, FeO_4^{2-}. The ionic half-equation for the conversion of ferrate(VI) to iron(III) is

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

(i) Ferrate(VI) decomposes in neutral or acid solution, but is stable in alkali.
Suggest why this is so.

(1)

.....

.....

.....

(ii) Write the equation for the disproportionation of iron(III) into iron(II) and ferrate(VI) in aqueous solution. State symbols are not required.

(2)

(iii) Use standard electrode potential values to determine the thermodynamic feasibility of this disproportionation.

(2)

(Total for question = 15 marks)

Q46. Which of the following will **not** reduce an acidified solution of potassium dichromate(VI)?

- ☒ **A** $(CH_3)_2C(OH)CH_3$
- ☒ **B** $FeSO_4$
- ☒ **C** $CH_3CH_2CH(OH)CH_3$
- ☒ **D** Zn

(Total for Question = 1 mark)

Q47.

Sodium hydride, NaH, can be used to generate hydrogen for fuel cells.

The sodium hydride is crushed in the presence of water to release the hydrogen gas for a fuel cell.

The overall equation for the reaction occurring in the fuel cell is

$$H_2(g) + ½O_2(g) \rightarrow H_2O(l)$$

In an alkaline fuel cell the oxidation half-equation is

$$H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$$

Deduce the reduction half-equation for the alkaline fuel cell.

State symbols are not required.

(1)

(Total for question = 1 mark)

Q48.

This question is about the $Ag^+(aq)|Ag(s)$ half-cell.

A student was asked to plan an experiment to measure the standard electrode potential of the $Ag^+(aq)|Ag(s)$ half-cell.

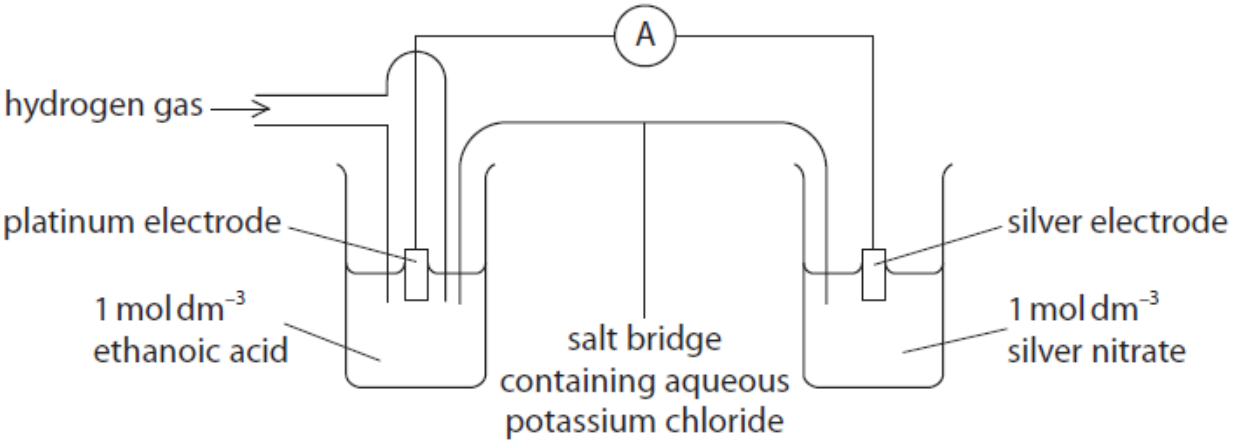
(i) State the conditions of temperature and pressure under which standard electrode potentials are measured.

(1)

.....

.....

(ii) The student drew the diagram shown.



Identify **three** mistakes in this diagram and the modifications that should be made to correct them.

(3)

Mistake in diagram	Modification needed to correct mistake

(Total for question = 4 marks)

Q49.

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.

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

The results of the titration are shown.

Run	Trial	1	2	3
Final volume / cm^3	17.50	34.10	17.20	34.10
Initial volume / cm^3	0.00	17.30	0.00	17.20
Titre / cm^3	17.50		17.20	
Concordant titres (✓)				
Mean titre / cm^3				

(i) Complete the table.

(2)

(ii) 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.

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

Use this equation and your mean titre from (i) to calculate the mass, in mg, of $KMnO_4$ in **one** tablet.

Give your answer to an appropriate number of significant figures.

(5)

(iii) A textbook suggested the conical flask should be heated during the titration, as the reaction between the MnO_4^- ions and the $C_2O_4^{2-}$ ions is slow.

Use these electrode potentials and your knowledge of homogeneous catalysis to deduce why the heating is very important at the start of the titration, but less important as the titration proceeds. Justify your answer.
You may include equations in your justification.

Electrode system	$E^{\ominus}$ / V
$2CO_2(g) + 2e^- \rightleftharpoons C_2O_4^{2-}(aq)$	+0.64
$Mn^{3+}(aq) + e^- \rightleftharpoons Mn^{2+}(aq)$	+1.49
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51

(4)

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

(Total for question = 11 marks)

Q50.

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.

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

Give the colour **change** at the end-point of the titration.

(1)

.....

(Total for question = 1 mark)

Q51.

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.

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

Give the reason why sulfuric acid was also added to the conical flask.

(1)

.....

.....

(Total for question = 1 mark)

Q52.

Brass is an alloy of copper and zinc.

A sample of brass was weighed on a balance, reading to two decimal places. The mass of the sample was recorded as 5.00 g.

This sample of brass was reacted with excess concentrated nitric acid and the resulting solution was made up to 250 ^{3} in a volumetric flask using distilled water.

25.0 cm^3 portions were taken from this solution using a pipette. Each portion was neutralised by adding sodium carbonate solution; and excess potassium iodide solution was then added. The liberated iodine was titrated with 0.250 mol dm^{-3} sodium thiosulfate solution, using a freshly prepared solution of starch as indicator.

The mean titre was 22.70 cm^3.

(i) Use the equation for the reaction to give two observations when nitric acid reacts with copper.

$$Cu(s) + 2NO_3^-(aq) + 4H^+(aq) \rightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$$

(2)

.....

.....

.....

.....

(ii) Complete the equation for the reaction between iodine and thiosulfate ions. Include state symbols.

(1)

$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow$

(iii) Copper(II) ions react with iodide ions to form iodine.

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

Calculate the percentage by mass of copper in the brass. Give your answer to an appropriate number of significant figures.

(5)

(iv) A student wants to identify the piece of apparatus that contributes most to measurement uncertainties in this experiment, so that the procedure can be modified.

The percentage measurement uncertainty is marked on the pipette as ±0.24%, and on the volumetric flask as ±0.08%.

By using appropriate calculations for the other apparatus used, deduce the **most** significant source of measurement uncertainty in this procedure.

(2)

.....

.....

.....

.....

.....

.....

Q53. Consider the following reaction.

$$S_2O_8^{2-}(aq) + 2I^{-}(aq) \rightarrow 2SO_4^{2-}(aq) + I_2(aq)$$

Which of the following ions could catalyse this reaction?

☒ **A** Zn^{2+}

☒ **B** Al^{3+}

☒ **C** Fe^{2+}

☒ **D** Na^{+}

(Total for Question = 1 mark)

Q54.

The value of E_{cell} indicates whether the cell reaction is thermodynamically feasible.
Which of the following is a correct statement about E_{cell}?

☒ **A** E_{cell} is directly proportional to the equilibrium constant.

☒ **B** E_{cell} is directly proportional to the entropy change of the system, ΔS_{system}.

☒ **C** E_{cell} is directly proportional to the total entropy change, ΔS_{total}.

☒ **D** The value of In E_{cell} is directly proportional to the total entropy change, ΔS_{total}.

(Total for question = 1 mark)

Q55.

(i) Complete the table below with the missing standard electrode potentials.
Use the table starting on page 14 of your Data Booklet.

(1)

Electrode system	Standard electrode potential $E^{\ominus}$ / V
$[VO^{2+}(aq) + 2H^{+}(aq)], [V^{3+}(aq) + H_2O(l)] \mid Pt$	
$[VO_2^{+}(aq) + 2H^{+}(aq)], [VO^{2+}(aq) + H_2O(l)] \mid Pt$	
$I_2(aq), 2I^{-}(aq) \mid Pt$	+0.54
$[2H^{+}(aq) + O_2(g)], [H_2O_2(aq)] \mid Pt$	+0.68

(ii) The colours of the different oxidation states of vanadium are shown below.

Oxidation state	Colour
+5	yellow
+4	blue
+3	green
+2	violet

For each of the following experiments, **A** and **B**, calculate the $E^{\ominus}$ value for the proposed reaction. Use your answers to predict whether or not a reaction occurs in each case.

Give the formula of the vanadium product formed where a reaction occurs and give **one** observation you would make in each experiment.

(6)

Experiment A: Hydrogen peroxide is added to an aqueous solution containing VO_2^+ ions.

.....

.....

.....

.....

.....

.....

.....

Experiment B: An aqueous solution of potassium iodide is added to an aqueous solution containing VO^{2+} ions.

.....

.....

.....

.....

.....

.....

.....

Q56. The electrode system based on the half-equation below has the standard electrode potential +1.51 V.

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

Which of the following statements about this electrode system is correct?

☒ **A** Changing the concentration of $Mn^{2+}(aq)$ would cause a change in the electrode potential.

☒ **B** $Mn^{2+}(aq)$ is acting as an oxidizing agent.

☒ **C** The electrode used is made of manganese.

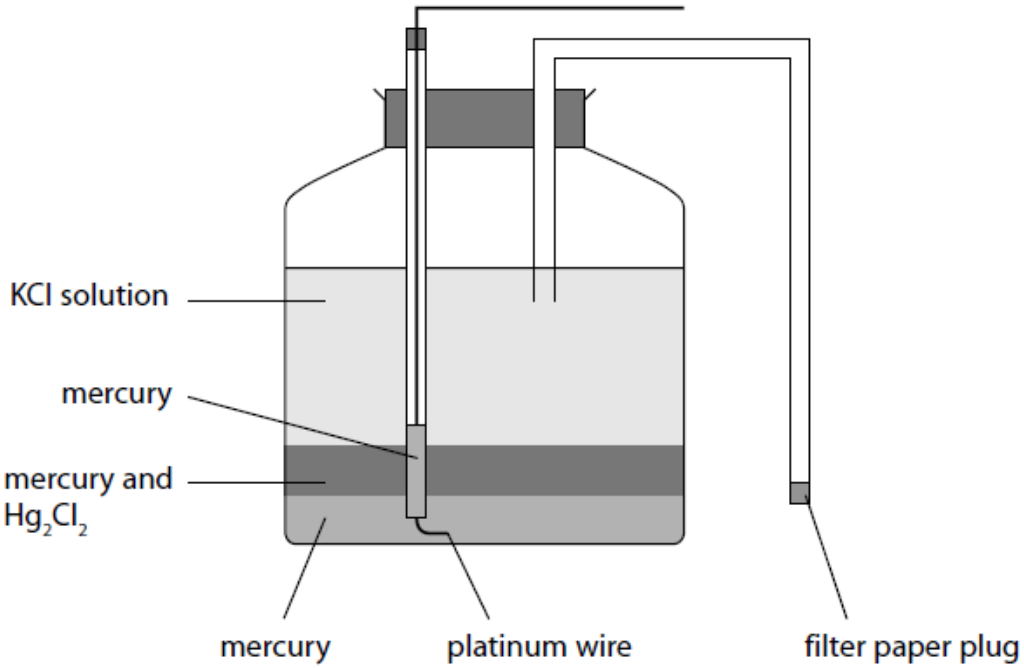
☒ **D** When connected to a standard hydrogen electrode, the resulting cell voltage would be

+0.51 V.

(Total for Question = 1 mark)

Q57.

The diagram shows a calomel electrode, which can be used for measuring standard electrode potentials.



(a) (i) State a solution that could be used in the side arm to act as the salt bridge.

(1)

.....

.....

(ii) Explain the purpose of the salt bridge.

(2)

.....

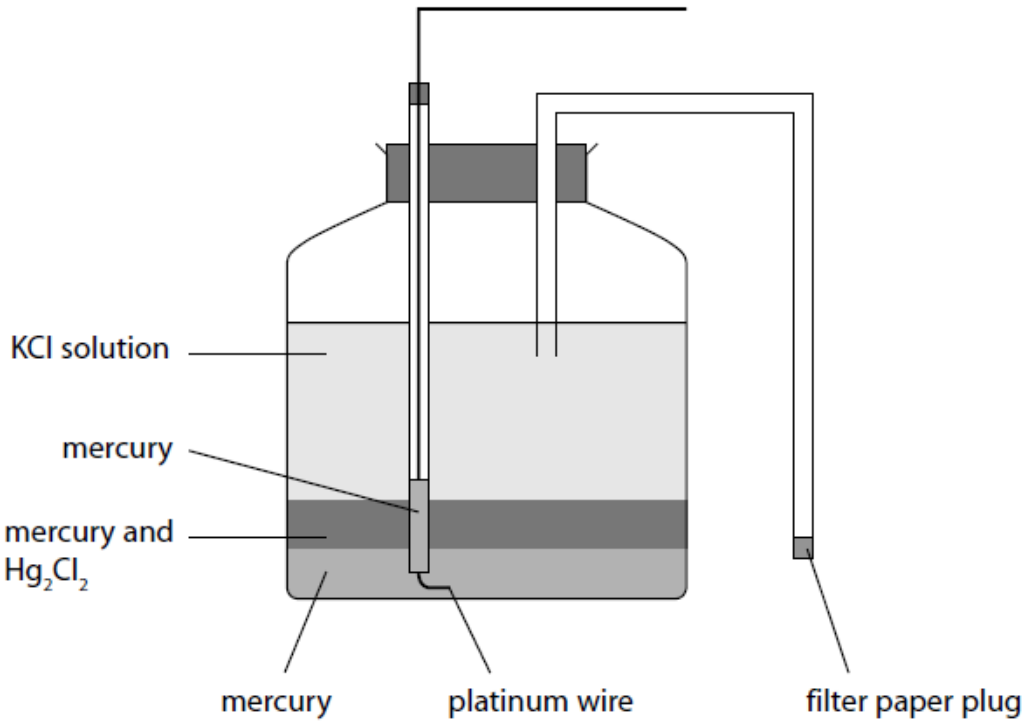
.....

.....

.....

(b) A calomel electrode can be used to measure the standard electromotive force (emf), $E^{\ominus}_{cell}$ of the cell:

$$Pt|[2Hg(l) + 2Cl^{-}(aq)],Hg_2Cl_2(s)::Ag^{+}(aq)|Ag(s)$$



(i) Complete the diagram above by drawing a suitable silver half-cell including the appropriate electric circuit.

(3)

(ii) State the concentration of the solution in the silver half-cell and a suitable experimental condition.

(2)

.....

.....

.....

.....

(c) (i) Use the Data Booklet to calculate the standard emf of the cell in b(i), given:

$$Pt|[2Hg(l) + 2Cl^-(aq)],Hg_2Cl_2(s) \quad E^\ominus = +0.27\ V$$

(2)

(ii) The emf of the cell shown below is:

$$Pt|[2Hg(l) + 2Cl^-(aq)],Hg_2Cl_2(s) :: [Fe^{3+}(aq),Fe^{2+}(aq)]|Pt \quad E^\ominus_{cell} = +0.50\ V$$

Use this value to calculate the standard emf of the following cell:

$$Pt[Fe^{2+}(aq),Fe^{3+}(aq)] :: Ag^+(aq)|Ag(s)$$

(1)

(iii) Write the overall equation for the cell reaction when current is being drawn.

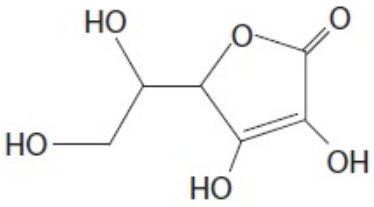
(1)

(iv) Calculate the numerical value of the equilibrium constant for this reaction at 298 K using $\Delta G^{\ominus} = -2892$ J mol^{-1} and the relationship $\Delta G^{\ominus} = -RT\ln K$.

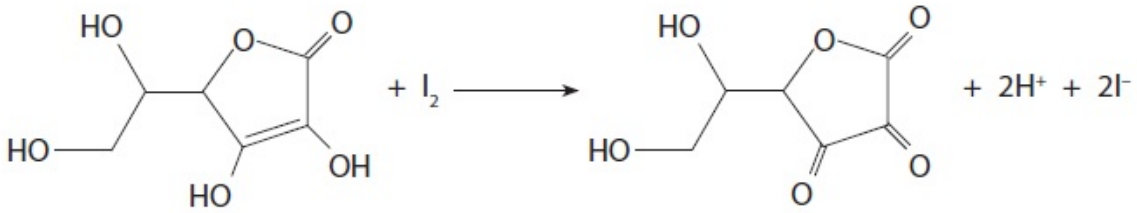
(2)

Q58.

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)

.....

.....

(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)

.....

.....

.....

.....

(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.

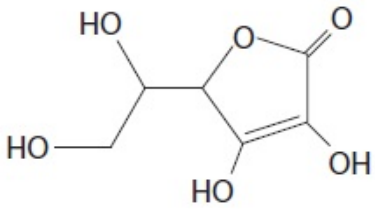
(1)

.....

.....

.....

(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)

.....

.....

.....

.....

(iii) Despite its name, ascorbic acid is not a carboxylic acid but it does contain an ester group.