

Please write clearly in block capitals.

Centre number

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Surname

*StemUp Model Solution*

Forename(s)

Candidate signature

# A-level CHEMISTRY

## Paper 1 Inorganic and Physical Chemistry

Tuesday 13 June 2017

Afternoon

Time allowed: 2 hours

### Materials

For this paper you must have:

- the Periodic Table/Data Booklet, provided as an insert (enclosed)
- a ruler with millimetre measurements
- a calculator, which you are expected to use where appropriate.

### Instructions

- Use black ink or black ball-point pen.
- Fill in the boxes at the top of the page.
- Answer **all** questions.
- You must answer the questions in the spaces provided.  
Do not write outside the box around each page or on blank pages.
- All working must be shown.
- Do all rough work in this book. Cross through any work you do not want to be marked.

### Information

- The marks for questions are shown in brackets.
- The maximum mark for this paper is 105.

| For Examiner's Use |      |
|--------------------|------|
| Question           | Mark |
| 1                  |      |
| 2                  |      |
| 3                  |      |
| 4                  |      |
| 5                  |      |
| 6                  |      |
| 7                  |      |
| 8                  |      |
| 9                  |      |
| 10                 |      |
| 11                 |      |
| <b>TOTAL</b>       |      |



J U N 1 7 7 4 0 5 1 0 1

Answer **all** questions in the spaces provided

0 1

This question is about silver iodide.

0 1 . 1

Define the term enthalpy of lattice formation.

[2 marks]

*The enthalpy change when 1 mole of a solid ionic compound (ionic lattice) is formed from its gaseous ions, e.g.  $\text{Ag}^+(\text{g}) + \text{I}^-(\text{g}) \rightarrow \text{AgI}(\text{s})$*

0 1 . 2

Some enthalpy change data are shown in **Table 1**.

**Table 1**

|                                                                                   | Enthalpy change<br>/ $\text{kJ mol}^{-1}$ |
|-----------------------------------------------------------------------------------|-------------------------------------------|
| $\text{AgI}(\text{s}) \rightarrow \text{Ag}^+(\text{aq}) + \text{I}^-(\text{aq})$ | +112                                      |
| $\text{Ag}^+(\text{g}) \rightarrow \text{Ag}^+(\text{aq})$                        | -464                                      |
| $\text{I}^-(\text{g}) \rightarrow \text{I}^-(\text{aq})$                          | -293                                      |

Use the data in **Table 1** to calculate the enthalpy of lattice formation of silver iodide.

[2 marks]



*Lattice dissociation enthalpy of AgI:*

$$\Delta H = (+112) + (+464) + (+293) = +869 \text{ kJ mol}^{-1}$$

$$\text{so } \Delta H(\text{lattice formation}) = -(+869)$$

$$= -869 \text{ kJ mol}^{-1}$$

Enthalpy of lattice formation -869  $\text{kJ mol}^{-1}$



0 1 . 3

A calculation of the enthalpy of lattice formation of silver iodide based on a perfect ionic model gives a smaller numerical value than the value calculated in Question 1.2

Explain this difference.

[2 marks]

*AgI is not perfectly ionic: the  $\text{Ag}^+$  ion polarises (distorts) the large  $\text{I}^-$  ion, so the bonding has some covalent character.*

*So the forces holding the lattice together are stronger than in the perfect ionic model, giving a more exothermic value.*

0 1 . 4

Identify a reagent that could be used to indicate the presence of iodide ions in an aqueous solution and describe the observation made.

[2 marks]

Reagent *silver nitrate solution,  $\text{AgNO}_3(\text{aq})$*

Observation *yellow precipitate (of AgI)*

8

Turn over for the next question



|   |   |
|---|---|
| 0 | 2 |
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This question is about acidic solutions.

|   |   |   |   |
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The acid dissociation constant,  $K_a$ , for ethanoic acid is given by the expression

$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]}$$

The value of  $K_a$  for ethanoic acid is  $1.74 \times 10^{-5} \text{ mol dm}^{-3}$  at  $25^\circ\text{C}$

A buffer solution with a pH of 3.87 was prepared using ethanoic acid and sodium ethanoate. In the buffer solution, the concentration of ethanoate ions was  $0.136 \text{ mol dm}^{-3}$

Calculate the concentration of the ethanoic acid in the buffer solution.  
Give your answer to three significant figures.

**[3 marks]**

$$[\text{H}^+] = 10^{-3.87} = 1.349 \times 10^{-4} \text{ mol dm}^{-3}$$

$$[\text{CH}_3\text{COOH}] = [\text{CH}_3\text{COO}^-][\text{H}^+] / K_a$$

$$= (0.136 \times 1.349 \times 10^{-4}) / (1.74 \times 10^{-5})$$

$$= 1.054 \text{ mol dm}^{-3}$$

Concentration of acid 1.05 mol dm<sup>-3</sup>



0 2 . 2

In a different buffer solution, the concentration of ethanoic acid was  $0.260 \text{ mol dm}^{-3}$  and the concentration of ethanoate ions was  $0.121 \text{ mol dm}^{-3}$

A  $7.00 \times 10^{-3} \text{ mol}$  sample of sodium hydroxide was added to  $500 \text{ cm}^3$  of this buffer solution.

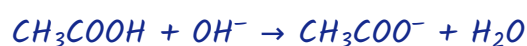
Calculate the pH of the buffer solution after the sodium hydroxide was added.  
Give your answer to two decimal places.

[6 marks]

*Moles in  $500 \text{ cm}^3$  before adding NaOH:*

$$n(\text{CH}_3\text{COOH}) = 0.260 \times 500/1000 = 0.130 \text{ mol}$$

$$n(\text{CH}_3\text{COO}^-) = 0.121 \times 500/1000 = 0.0605 \text{ mol}$$



$$\text{After: } n(\text{CH}_3\text{COOH}) = 0.130 - 0.00700 = 0.123 \text{ mol}$$

$$n(\text{CH}_3\text{COO}^-) = 0.0605 + 0.00700 = 0.0675 \text{ mol}$$

$$[\text{H}^+] = K_a \times [\text{CH}_3\text{COOH}] / [\text{CH}_3\text{COO}^-]$$

*(same volume, so the ratio of moles can be used)*

$$[\text{H}^+] = 1.74 \times 10^{-5} \times 0.123 / 0.0675$$

$$= 3.171 \times 10^{-5} \text{ mol dm}^{-3}$$

$$\text{pH} = -\log(3.171 \times 10^{-5}) = 4.50$$

pH of buffer solution 4.50



0 3

The ionic product of water,  $K_w = 2.93 \times 10^{-15} \text{ mol}^2 \text{ dm}^{-6}$  at  $10^\circ \text{C}$

0 3 . 1

Which is the correct expression for  $K_w$ ?  
Tick (✓) **one** box.

[1 mark]

A  $K_w = \frac{[\text{H}_2\text{O}]}{[\text{H}^+][\text{OH}^-]}$  ☐

B  $K_w = [\text{H}^+][\text{H}_2\text{O}]$  ☐

C  $K_w = [\text{H}^+][\text{OH}^-]$  ☒

D  $K_w = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}$  ☐

0 3 . 2

Calculate the pH of pure water at  $10^\circ \text{C}$   
Give your answer to two decimal places.

[2 marks]

$$[\text{H}^+] = \sqrt{K_w} = \sqrt{(2.93 \times 10^{-15})}$$

$$= 5.41 \times 10^{-8} \text{ mol dm}^{-3}$$

$$\text{pH} = -\log(5.41 \times 10^{-8}) = 7.27$$

pH of water 7.27

0 3 . 3

Suggest why this pure water at  $10^\circ \text{C}$  is **not** alkaline.

[1 mark]

In pure water  $[\text{H}^+] = [\text{OH}^-]$  (equal numbers of  $\text{H}^+$  and  $\text{OH}^-$   
ions), so it is neutral, not alkaline.



0 3 . 4

Calculate the pH of a  $0.0131 \text{ mol dm}^{-3}$  solution of calcium hydroxide at  $10^\circ\text{C}$ .  
Give your answer to two decimal places.

[3 marks]



$$[\text{OH}^-] = 2 \times 0.0131 = 0.0262 \text{ mol dm}^{-3}$$

$$[\text{H}^+] = K_w / [\text{OH}^-] = 2.93 \times 10^{-15} / 0.0262$$

$$= 1.118 \times 10^{-13} \text{ mol dm}^{-3}$$

$$\text{pH} = -\log(1.118 \times 10^{-13}) = 12.95$$

pH of solution 12.95

0 3 . 5

The  $0.0131 \text{ mol dm}^{-3}$  calcium hydroxide solution at  $10^\circ\text{C}$  was a saturated solution.

A student added  $0.0131 \text{ mol}$  of magnesium hydroxide to  $1.00 \text{ dm}^3$  of water at  $10^\circ\text{C}$  and stirred the mixture until no more solid dissolved.

Predict whether the pH of the magnesium hydroxide solution formed at  $10^\circ\text{C}$  is larger than, smaller than or the same as the pH of the calcium hydroxide solution at  $10^\circ\text{C}$ .

Explain your answer.

[2 marks]

pH of magnesium hydroxide compared to calcium hydroxide

Smaller (lower pH, less alkaline)

Explanation Magnesium hydroxide is less soluble than calcium

hydroxide (solubility of the hydroxides increases down Group 2), so

fewer  $\text{OH}^-$  ions dissolve and  $[\text{OH}^-]$  is lower.



0 4

A sample of titanium was ionised by electron impact in a time of flight (TOF) mass spectrometer. Information from the mass spectrum about the isotopes of titanium in the sample is shown in **Table 2**.

**Table 2**

|               |     |     |      |     |
|---------------|-----|-----|------|-----|
| m/z           | 46  | 47  | 48   | 49  |
| Abundance / % | 9.1 | 7.8 | 74.6 | 8.5 |

0 4 . 1

Calculate the relative atomic mass of titanium in this sample.  
Give your answer to one decimal place.

**[2 marks]**

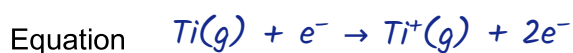
$$A_r = (46 \times 9.1 + 47 \times 7.8 + 48 \times 74.6 + 49 \times 8.5) / 100$$

$$= 4782.5 / 100 = 47.825$$

Relative atomic mass of titanium in this sample 47.8

0 4 . 2

Write an equation, including state symbols, to show how an atom of titanium is ionised by electron impact and give the m/z value of the ion that would reach the detector first.

**[2 marks]**

m/z value 46

0 4 . 3

Calculate the mass, in kg, of one atom of  $^{49}Ti$

The Avogadro constant  $L = 6.022 \times 10^{23} \text{ mol}^{-1}$

**[1 mark]**

$$\text{mass of 1 mol of } ^{49}Ti = 49 \text{ g} = 0.049 \text{ kg}$$

$$\text{mass of one atom} = 0.049 / 6.022 \times 10^{23} = 8.137 \times 10^{-26} \text{ kg}$$

Mass  $8.137 \times 10^{-26}$  kg



0 4 . 4

In a TOF mass spectrometer the time of flight,  $t$ , of an ion is shown by the equation

$$t = d \sqrt{\frac{m}{2E}}$$

In this equation  $d$  is the length of the flight tube,  $m$  is the mass, in kg, of an ion and  $E$  is the kinetic energy of the ions.

In this spectrometer, the kinetic energy of an ion in the flight tube is  $1.013 \times 10^{-13} \text{ J}$

The time of flight of a  $^{49}\text{Ti}^+$  ion is  $9.816 \times 10^{-7} \text{ s}$

Calculate the time of flight of the  $^{47}\text{Ti}^+$  ion.  
Give your answer to the appropriate number of significant figures.

[3 marks]

*Rearranging:  $d = t \sqrt{(2E / m)}$*

*$d$  and  $E$  are the same for both ions, so  $t$  is proportional to  $\sqrt{m}$ :*

$$t_{47} / \sqrt{47} = t_{49} / \sqrt{49}$$

$$t_{47} = t_{49} \times \sqrt{(47 / 49)}$$

$$= 9.816 \times 10^{-7} \times \sqrt{(47 / 49)}$$

$$= 9.816 \times 10^{-7} \times 0.97938$$

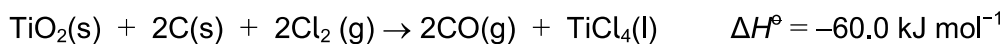
$$= 9.614 \times 10^{-7} \text{ s (4 s.f., as in the data)}$$

Time of flight  $9.614 \times 10^{-7}$  s



0 5

Titanium(IV) chloride can be made from titanium(IV) oxide as shown in the equation.



0 5 . 1

Some entropy data are shown in **Table 3**.

**Table 3**

| Substance                                | TiO <sub>2</sub> (s) | C(s) | Cl <sub>2</sub> (g) | CO(g) | TiCl <sub>4</sub> (l) |
|------------------------------------------|----------------------|------|---------------------|-------|-----------------------|
| S° / J K <sup>-1</sup> mol <sup>-1</sup> | 50.2                 | 5.70 | 223                 | 198   | 253                   |

Use the equation and the data in **Table 3** to calculate the Gibbs free-energy change for this reaction at 989 °C

Give your answer to the appropriate number of significant figures.

Use your answer to explain whether this reaction is feasible.

**[6 marks]**

$$\Delta S = \Sigma S(\text{products}) - \Sigma S(\text{reactants})$$

$$= (2 \times 198 + 253) - (50.2 + 2 \times 5.70 + 2 \times 223)$$

$$= 649 - 507.6 = +141.4 \text{ J K}^{-1} \text{ mol}^{-1}$$

$$T = 989 + 273 = 1262 \text{ K}$$

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

$$= -60.0 - (1262 \times 141.4 \times 10^{-3})$$

$$= -60.0 - 178.4$$

$$= -238 \text{ kJ mol}^{-1} \text{ (3 s.f.)}$$

Gibbs free-energy change -238 kJ mol<sup>-1</sup>

Explanation The reaction is feasible at 989 °C because ΔG is  
negative (less than zero).



**Turn over for the next question**

**DO NOT WRITE ON THIS PAGE  
ANSWER IN THE SPACES PROVIDED**



0 6

This question is about some Period 3 elements and their oxides.

0 6 . 1

Write an equation for the reaction of phosphorus with an excess of oxygen.

[1 mark]



0 6 . 2

Describe a test you could carry out in a test tube to distinguish between sodium oxide and the product of the reaction in Question 6.1

[3 marks]

Add a little water to each oxide in separate test tubes and shake,  
then add a few drops of universal indicator to each solution.

Sodium oxide forms  $NaOH(aq)$ : the indicator turns blue / purple

(pH about 13 to 14).

$P_4O_{10}$  forms  $H_3PO_4(aq)$ : the indicator turns red (pH about 0 to 1).

0 6 . 3

State the type of crystal structure shown in silicon dioxide and in sulfur trioxide.

[2 marks]

Silicon dioxide giant covalent (macromolecular)

Sulfur trioxide molecular (simple molecules)

0 6 . 4

Explain why silicon dioxide has a higher melting point than sulfur trioxide.

[4 marks]

Silicon dioxide is giant covalent, so to melt it many strong  
covalent bonds between the atoms must be broken.

Sulfur trioxide is molecular, so to melt it only the weak van der  
Waals forces between the molecules must be overcome; no  
covalent bonds are broken.

Covalent bonds are much stronger than van der Waals forces, so  
much more energy is needed to break them, so  $SiO_2$  has the  
higher melting point.



0 6 . 5

Write an equation for the reaction of sulfur trioxide with potassium hydroxide solution.

[1 mark]



0 6 . 6

Write an equation for the reaction of an excess of magnesium oxide with phosphoric acid.

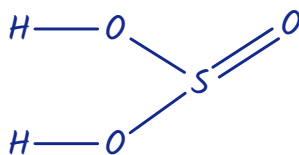
[1 mark]



0 6 . 7

Draw the displayed formula of the undissociated acid formed when sulfur dioxide reacts with water.

[1 mark]



13

Turn over for the next question



0 7

Solution **A** contains the compound  $[\text{Cu}(\text{H}_2\text{O})_6]\text{Cl}_2$ 

0 7 . 1

State the type of bonding between the oxygen and hydrogen in this compound.

[1 mark]

*Covalent (bonds)*

0 7 . 2

State why the chloride ions in this compound are **not** considered to be ligands.

[1 mark]

*The chloride ions do not donate a lone pair of electrons to the  $\text{Cu}^{2+}$  ion, so they do not form coordinate bonds; they are only held by ionic attraction.*

0 7 . 3

An excess of ammonia was added to a sample of solution **A** to form solution **B**.Write an ionic equation for the reaction that occurs when solution **A** is converted into solution **B** and state the colour of solution **B**.

[2 marks]

Equation  *$[\text{Cu}(\text{H}_2\text{O})_6]^{2+} + 4\text{NH}_3 \rightarrow [\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+} + 4\text{H}_2\text{O}$* Colour *deep blue (solution)*

0 7 . 4

Aqueous sodium carbonate was added to another sample of solution **A** to form a blue-green solid **C**.Identify the blue-green solid **C**.

[1 mark]

*copper(II) carbonate,  $\text{CuCO}_3$* 

0 7 . 5

Reagent **D** was added to another sample of solution **A** to form a yellow-green solution.Identify reagent **D** and write an ionic equation for the reaction that occurs when the yellow-green solution is formed from solution **A**.

[2 marks]

Identity of reagent **D** *concentrated hydrochloric acid,  $\text{HCl}$* Equation  *$[\text{Cu}(\text{H}_2\text{O})_6]^{2+} + 4\text{Cl}^- \rightarrow [\text{CuCl}_4]^{2-} + 6\text{H}_2\text{O}$* 

07 . 6

Explain why colorimetry cannot be used to determine the concentration of solutions containing  $[\text{CuCl}_2]^-$

In your answer refer to the electron configuration of the metal ion.

[2 marks]

*In  $[\text{CuCl}_2]^-$  copper is  $\text{Cu}^+$ ,  $[\text{Ar}] 3d^0$ , so its 3d sub-shell is full.*

*No d-electron transitions can occur, so the ion cannot absorb*

*visible light; the solution is colourless, so colorimetry cannot be*

*used.*

9

Turn over for the next question



0 8

This question is about ion testing.

0 8 . 1

Describe how a student could distinguish between aqueous solutions of potassium nitrate,  $\text{KNO}_3$ , and potassium sulfate,  $\text{K}_2\text{SO}_4$ , using **one** simple test-tube reaction.

**[3 marks]**

Reagent barium chloride solution,  $\text{BaCl}_2(\text{aq})$

Observation with  $\text{KNO}_3(\text{aq})$  no visible change (colourless solution, no precipitate)

Observation with  $\text{K}_2\text{SO}_4(\text{aq})$  white precipitate forms ( $\text{BaSO}_4$ )

0 8 . 2

Describe how a student could distinguish between aqueous solutions of magnesium chloride,  $\text{MgCl}_2$ , and aluminium chloride,  $\text{AlCl}_3$ , using **one** simple test-tube reaction.

**[3 marks]**

Reagent sodium hydroxide solution,  $\text{NaOH}(\text{aq})$

Observation with  $\text{MgCl}_2(\text{aq})$  white precipitate which does not dissolve in excess  $\text{NaOH}$

Observation with  $\text{AlCl}_3(\text{aq})$  white precipitate which dissolves in excess  $\text{NaOH}$ , giving a colourless solution

6

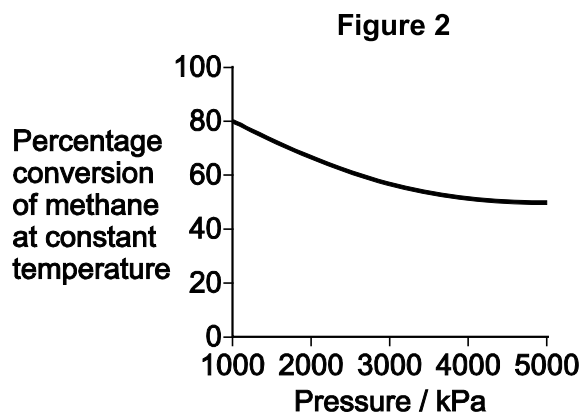
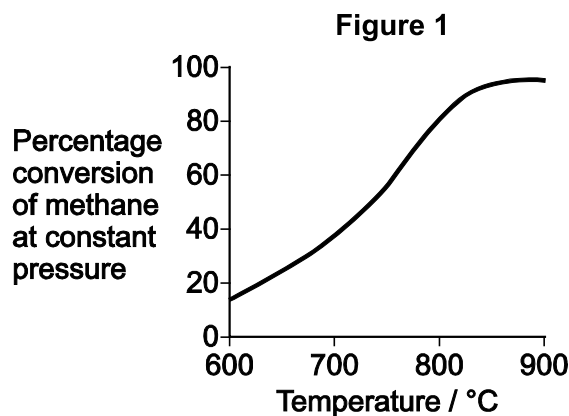


0 9

There are several stages in the industrial production of methanol from methane.

0 9 . 1

The first stage involves a gaseous equilibrium between the reactants (methane and steam), and some gaseous products. **Figures 1** and **2** show the percentage conversion of methane into the gaseous products under different conditions at equilibrium.



Deduce the optimum conditions for the industrial conversion of methane and steam into the gaseous products.

Explain your deductions.

[6 marks]

*Figure 1: at constant pressure the percentage conversion of methane increases as the temperature increases, so the forward reaction is endothermic.*

*However, above about 850 °C the conversion levels off at about 95%, so increasing the temperature further gives no significant increase in yield.*

*Figure 2: at constant temperature the percentage conversion decreases as the pressure increases, because there are fewer moles of gas on the left (2 mol) than on the right, so the equilibrium shifts to the left at high pressure.*

*Optimum temperature: about 850 °C (in the range 780 to 880 °C).*

*A high temperature gives a high yield and a fast rate, but high temperatures have high energy (fuel) costs and are expensive. Above*



about 850 °C the yield no longer increases, so there is no gain from using a higher temperature.

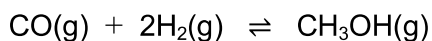
Optimum pressure: a moderate pressure of about 1000 to 2000 kPa. Although a low pressure gives the highest yield, a very low pressure gives too slow a rate (fewer collisions per second), while a higher pressure lowers the yield and is expensive to produce and needs stronger equipment.

So a compromise pressure of about 1000 to 2000 kPa is used together with a temperature of about 850 °C, giving a high conversion at an acceptable rate and cost.



0 9 . 2

The equation shows the final stage in the production of methanol.



20.1 mol of carbon monoxide and 24.2 mol of hydrogen were placed in a sealed container. An equilibrium was established at 600 K. The equilibrium mixture contained 2.16 mol of methanol.

Calculate the amount, in moles, of carbon monoxide and of hydrogen in the equilibrium mixture.

[2 marks]

*CO + 2H<sub>2</sub> ⇌ CH<sub>3</sub>OH: forming 2.16 mol CH<sub>3</sub>OH uses*

*2.16 mol CO and 2 × 2.16 = 4.32 mol H<sub>2</sub>*

*n(CO) = 20.1 – 2.16 = 17.94 mol*

*n(H<sub>2</sub>) = 24.2 – 4.32 = 19.88 mol*

Amount of carbon monoxide 17.9 mol

Amount of hydrogen 19.9 mol

**Question 9 continues on the next page**



09 . 3

A different mixture of carbon monoxide and hydrogen was allowed to reach equilibrium at 600 K

At equilibrium, the mixture contained 2.76 mol of carbon monoxide, 4.51 mol of hydrogen and 0.360 mol of methanol. The total pressure was 630 kPa

Calculate a value for the equilibrium constant,  $K_p$ , for this reaction at 600 K and state its units.

[6 marks]

$$K_p = p(\text{CH}_3\text{OH}) / (p(\text{CO}) \times p(\text{H}_2)^2)$$

$$\text{Total moles} = 2.76 + 4.51 + 0.360 = 7.63 \text{ mol}$$

$$p(\text{CO}) = (2.76 / 7.63) \times 630 = 227.9 \text{ kPa}$$

$$p(\text{H}_2) = (4.51 / 7.63) \times 630 = 372.4 \text{ kPa}$$

$$p(\text{CH}_3\text{OH}) = (0.360 / 7.63) \times 630 = 29.72 \text{ kPa}$$

$$K_p = 29.72 / (227.9 \times 372.4^2)$$

$$= 9.41 \times 10^{-7} \text{ kPa}^{-2}$$

Value of  $K_p$   $9.41 \times 10^{-7}$  Units  $\text{kPa}^{-2}$



1 0

**Table 4** shows some electrode half-equations and their standard electrode potentials.

**Table 4**

| Electrode half-equation                                                                                                                 | $E^\circ / \text{V}$ |
|-----------------------------------------------------------------------------------------------------------------------------------------|----------------------|
| $\text{Cl}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{Cl}^-(\text{aq})$                                                               | +1.36                |
| $\text{NO}_3^-(\text{aq}) + 4\text{H}^+(\text{aq}) + 3\text{e}^- \rightarrow \text{NO}(\text{aq}) + 2\text{H}_2\text{O}(\text{aq})$     | +0.96                |
| $\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq})$                                                          | +0.77                |
| $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$                                                               | +0.34                |
| $\text{SO}_4^{2-}(\text{aq}) + 4\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{SO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{aq})$ | +0.17                |
| $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$                                                                 | 0.00                 |
| $\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Fe}(\text{s})$                                                               | -0.44                |

1 0 . 1

Deduce the oxidation state of nitrogen in  $\text{NO}_3^-$  and in NO

**[2 marks]**

Nitrogen in  $\text{NO}_3^-$  +5

Nitrogen in NO +2

1 0 . 2

State the weakest reducing agent in **Table 4**.

**[1 mark]**

$\text{Cl}^-$  (chloride ions)

1 0 . 3

Write the conventional representation of the cell that has an EMF of +0.43 V

**[2 marks]**

$\text{Cu}(\text{s}) \mid \text{Cu}^{2+}(\text{aq}) \parallel \text{Fe}^{3+}(\text{aq}), \text{Fe}^{2+}(\text{aq}) \mid \text{Pt}(\text{s})$



1 0 . 4

Use data from **Table 4** to identify an acid that will oxidise copper.

Explain your choice of acid.

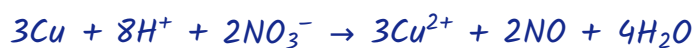
Use these data to suggest a possible equation for the reaction.

Calculate the EMF of the cell that has the same overall reaction.

[4 marks]

*Nitric acid,  $\text{HNO}_3$*

*$E^\ominus(\text{NO}_3^-/\text{NO}) = +0.96 \text{ V}$  is more positive than  $E^\ominus(\text{Cu}^{2+}/\text{Cu}) = +0.34 \text{ V}$ , so  $\text{NO}_3^-$  is a better oxidising agent than  $\text{Cu}^{2+}$  and can oxidise Cu to  $\text{Cu}^{2+}$ .*



$$\text{EMF} = 0.96 - 0.34 = +0.62 \text{ V}$$



1 1

This question is about compounds containing ethanedioate ions.

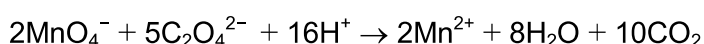
1 1 . 1

A white solid is a mixture of sodium ethanedioate ( $\text{Na}_2\text{C}_2\text{O}_4$ ), ethanedioic acid dihydrate ( $\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$ ) and an inert solid. A volumetric flask contained 1.90 g of this solid mixture in  $250 \text{ cm}^3$  of aqueous solution.

Two different titrations were carried out using this solution.

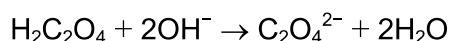
In the first titration  $25.0 \text{ cm}^3$  of the solution were added to an excess of sulfuric acid in a conical flask. The flask and contents were heated to  $60^\circ\text{C}$  and then titrated with a  $0.0200 \text{ mol dm}^{-3}$  solution of potassium manganate(VII). When  $26.50 \text{ cm}^3$  of potassium manganate(VII) had been added the solution changed colour.

The equation for this reaction is



In the second titration  $25.0 \text{ cm}^3$  of the solution were titrated with a  $0.100 \text{ mol dm}^{-3}$  solution of sodium hydroxide using phenolphthalein as an indicator. The indicator changed colour after the addition of  $10.45 \text{ cm}^3$  of sodium hydroxide solution.

The equation for this reaction is



Calculate the percentage by mass of sodium ethanedioate in the white solid.

Give your answer to the appropriate number of significant figures.  
Show your working.

[8 marks]

*Titration 1 (all the ethanedioate, from acid and salt):*

$$n(\text{MnO}_4^-) = 26.50 \times 0.0200 / 1000 = 5.30 \times 10^{-4} \text{ mol}$$

$$n(\text{C}_2\text{O}_4^{2-}) \text{ in } 25.0 \text{ cm}^3 = 5.30 \times 10^{-4} \times 5/2$$

$$= 1.325 \times 10^{-3} \text{ mol}$$

*Titration 2 (only the acid reacts):*

$$n(\text{NaOH}) = 10.45 \times 0.100 / 1000 = 1.045 \times 10^{-3} \text{ mol}$$

$$n(\text{H}_2\text{C}_2\text{O}_4) \text{ in } 25.0 \text{ cm}^3 = 1.045 \times 10^{-3} \div 2$$

$$= 5.225 \times 10^{-4} \text{ mol}$$



$$n(\text{Na}_2\text{C}_2\text{O}_4) \text{ in } 25.0 \text{ cm}^3 = 1.325 \times 10^{-3} - 5.225 \times 10^{-4}$$

$$= 8.025 \times 10^{-4} \text{ mol}$$

$$n(\text{Na}_2\text{C}_2\text{O}_4) \text{ in } 250 \text{ cm}^3 = 8.025 \times 10^{-4} \times 10$$

$$= 8.025 \times 10^{-3} \text{ mol}$$

$$M_r(\text{Na}_2\text{C}_2\text{O}_4) = 2(23.0) + 2(12.0) + 4(16.0) = 134.0$$

$$\text{mass of Na}_2\text{C}_2\text{O}_4 = 8.025 \times 10^{-3} \times 134.0 = 1.075 \text{ g}$$

$$\% \text{ by mass} = (1.075 / 1.90) \times 100$$

$$= 56.6 \% \text{ (3 s.f.)}$$

Percentage by mass of sodium ethanedioate 56.6 %



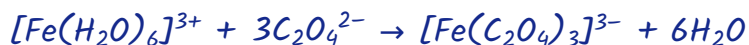
1 1 . 2

Ethanedioate ions react with aqueous iron(III) ions in a ligand substitution reaction.

Write an equation for this reaction.

Suggest why the value of the enthalpy change for this reaction is close to zero.

[2 marks]



Six Fe-O bonds are broken and six Fe-O bonds are made, the same number and type of bond, so  $\Delta H$  is close to zero.

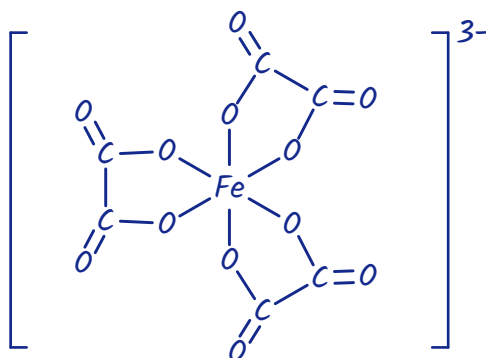
1 1 . 3

Draw the displayed formula of the iron complex produced in the reaction in Question 11.2

Indicate the value of the O—Fe—O bond angle.

State the type of isomerism shown by the iron complex.

[3 marks]



Bond angle 90°

Type of isomerism optical

1 1 . 4

Ethanedioate ions are poisonous because they react with iron ions in the body. Ethanedioate ions are present in foods such as broccoli and spinach.

Suggest one reason why people who eat these foods do not suffer from poisoning.

[1 mark]

Ethanedioate ions are only present in small amounts (low concentration) in these foods, so too little is eaten to be harmful.

END OF QUESTIONS



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