Anderson Serangoon Junior College 2020 JC2 H2 Chemistry Paper 4 Suggested Mark Scheme

1 Determination of the oxidation number of iodine in a compound

In this experiment you will determine the oxidation number of iodine in one of its compounds by titration.

You are provided with the following:

FA 1 is a 0.0197 mol dm⁻³ solution of the iodine–containing compound. **FA 2** is dilute sulfuric acid, H_2SO_4 . **FA 3** is aqueous potassium iodide, KI. **FA 4** is 0.105 mol dm⁻³ sodium thiosulfate, Na₂S₂O₃. starch indicator

FA 1 reacts with excess acidified potassium iodide to produce iodine, I₂. This iodine is then titrated with aqueous sodium thiosulfate using starch indicator.

(a) Titration of iodine formed against FA 4

- 1. Fill the burette with FA 4. (You will use the same burette for Question 2.)
- 2. Pipette 25.0 cm³ of **FA 1** into a conical flask.
- 3. Using a 25 cm³ measuring cylinder, add 10 cm³ of **FA 2** to the same conical flask.
- 4. Using the same measuring cylinder, add 20 cm³ of **FA 3** to the mixture in the conical flask. The mixture will now be a red–brown colour, due to iodine produced.
- 5. Add **FA 4** from the burette until the mixture becomes light brown.
- 6. Then add 10 drops of starch indicator. The mixture will change to a dark blue colour.
- 7. Continue titrating until the mixture becomes colourless. This is the end-point.
- 8. Carry out as many accurate titrations as you think necessary to obtain consistent results.
- 9. Record in a suitable form below all of your burette readings and the volume of **FA 4** added in each accurate titration.

(i) Results

	1	2
initial burette reading / cm ³	0.00	10.00
final burette reading / cm ³	28.40	38.25
volume of FA 4 / cm ³	28.40	28.35

[5]

(ii) From your accurate titration results, obtain a suitable value for the volume of **FA 4** to be used in your calculations. Show clearly how you obtained this value.

$$\frac{28.35 + 28.40}{2} = 28.38 \text{ cm}^3$$

The iodine produced required 28.38 cm³ of FA 4. [1]

- (b) Show your working and appropriate significant figures in the final answer to each step of your calculations.
 - (i) Calculate the number of moles of sodium thiosulfate in the volume of FA 4 calculated in (a)(ii).

 $n(S_2O_3^{2-})$ used = 0.105 x $\frac{28.38}{1000}$ mol = 2.98 x 10⁻³ mol

moles of $Na_2S_2O_3 =$ [1]

(ii) The equation for the reaction of iodine with sodium thiosulfate is shown.

 $I_2(aq) + 2Na_2S_2O_3(aq) \rightarrow Na_2S_4O_6(aq) + 2NaI(aq)$

Calculate the number of moles of iodine that reacted with the sodium thiosulfate calculated in (i).

 $n(I_2)$ reacted = $\frac{1}{2} \times 2.98 \times 10^{-3}$ = 1.49 x 10⁻³ mol

(iii) Use the information on page 2 to calculate the number of moles of iodinecontaining compound in the 25 cm³ of **FA1** used in each titration.

(iv) Use your answers to (b)(ii) and (b)(iii) to calculate the number of moles of iodine produced when 1 mole of the iodine-containing compound in FA 1 reacts with excess FA 3.

Give your answer as an integer.

no. of moles =
$$\frac{1.49 \times 10^{-3}}{4.93 \times 10^{-4}}$$

= 3

moles of $I_2 =$ [1]

(v) The anion in **FA 1** is IO_x^- where x is the number of oxygen atoms present in the formula.

Use your answer to (iv) to balance the ionic equation for the reaction between **FA 1** and **FA 3** under acidic conditions. Hence deduce the value of *x* in the formula IO_x^{-1} .

 $IO...^{-} + \dots I^{-} + \dots H^{+} \rightarrow \dots I_{2} + \dots H_{2}O$

By observation:	$IO_{}^{-} + \ldots I^{-} + \ldots H^{+} \rightarrow 3I_{2} + \ldots H_{2}O$	(use ans from (iv))
	$IO_{}^{-}$ + 5 I^{-} + H^{+} \rightarrow $3I_{2}$ + $H_{2}O$	(balance I)
	$IO_{}^{-}$ + 5I ⁻ + 6H ⁺ \rightarrow 3I ₂ +H ₂ O	(balance charge)
	$IO_{}^{-}$ + 5I ⁻ + 6H ⁺ \rightarrow 3I ₂ + 3 H ₂ O	(balance H)
	IO_3^- + $5I^-$ + $6H^+ \rightarrow 3I_2$ + $3H_2O$	(balance O)
therefore $x = 3$		

x =[2]

(vi) Calculate the oxidation state of iodine in FA 1. (If you were unable to calculate x in part (v), assume that x = 4.) Let the O.S. of iodine in FA 1 be a.

a + 3(-2) = -1 a = +5

oxidation state of iodine =[2]

[Total: 14]

2 To investigate the effect of concentration changes on the rate of a reaction.

FA 5 is 0.0500 mol dm⁻³ acidified iron(III) chloride, FeC l_3 .

FA 6 is 0.0500 mol dm⁻³ potassium iodide, KI.

Iron(III) ions oxidise iodide ions, I^- , to iodine, I_2 as shown in equation 1.

equation 1 $2Fe^{3+}(aq) + 2I^{-}(aq) \rightarrow 2Fe^{2+}(aq) + I_2(aq)$

In this experiment you will investigate how the rate of this reaction is affected by the concentration of Fe³⁺ ions. To do this you will add thiosulfate ions, $S_2O_3^{2-}$, and starch indicator to a mixture of Fe³⁺(aq) and I⁻(aq). The iodine produced by the reaction reacts immediately with the thiosulfate ions and is reduced back to iodide as shown in equation 2.

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

When all the thiosulfate has reacted, the iodine remaining in solution turns the starch indicator blue–black. The rate of reaction can be determined by timing how long it takes for the reaction mixture to turn blue–black.

(a) Preparation of FA 7 by dilution of FA 4

FA 7 which contains 0.00500 mol dm⁻³ of sodium thiosulfate, Na₂S₂O₃, can be prepared by diluting **FA 4** from **Question 1** in a 250 cm³ graduated flask.

(i) The concentration of sodium thiosulfate, $Na_2S_2O_3$, in **FA 4** is 0.105 mol dm⁻³.

Show, by calculation, that the volume of **FA 4** needed to prepare this diluted solution in the 250 cm³ graduated flask is 11.90 cm³. [1]

Let V_1 cm³ be the volume of **FA 4** needed for dilution.

 $c_1V_1 = c_2V_2$ 0.105 x V₁ = 0.005 x 250 V₁ = 11.90 cm³

(ii) Using the burette from **Question 1**, measure 11.90 cm³ of **FA 4** into a 250 cm³ graduated flask labelled **FA 7**. Make up to the mark with distilled water.

Record the volume of FA 4 added to the flask in the space provided.

volume of **FA 4** added = cm³

(b) Experiment 1

- 1. Fill the burette labelled **FA 5** with **FA 5**.
- 2. Run 20.00 cm³ of **FA 5** into a 100 cm³ beaker.
- 3. Using the **same** measuring cylinder, add the following to the **second** 100 cm³ beaker:
 - 10 cm³ of **FA 6**
 - 20 cm³ of **FA 7**
 - 10 cm³ of starch solution
- 4. Add the contents of the first beaker to the second beaker and start timing immediately.
- 5. Stir the mixture once and place the beaker on a white tile.
- 6. Stop timing as soon as the solution turns blue–black. Ignore any colour changes that occur before the intense blue–black colouration.
- 7. Record this reaction time, to the nearest second, in the space provided on page 8.
- 8. Rinse both beakers and shake dry. Rinse and dry the glass rod.

Experiment 2

- 1. Run 10.00 cm³ of **FA 5** into a 100 cm³ beaker.
- 2. Using another measuring cylinder, add 10.0 cm³ of distilled water into the beaker containing **FA 5**.
- 3. Using the measuring cylinder from Experiment 1, add the following to the **second** 100 cm³ beaker:
 - 10 cm³ of **FA 6**
 - 20 cm³ of **FA7**
 - 10 cm³ of starch solution
- 4. Add the contents of the first beaker to the second beaker and start timing immediately.
- 5. Stir the mixture once and place the beaker on a white tile.
- 6. Stop timing as soon as the solution turns blue–black. Ignore any colour changes that occur before the intense blue–black colouration.
- 7. Record this reaction time, to the nearest second, in the space provided on page 8.
- 1. Rinse both beakers and shake dry. Rinse and dry the glass rod.

Experiment 3 to 5

Carry out **three** further experiments to investigate how the reaction time changes with different volumes of **FA 5**.

Remember that the combined volume of **FA 5** and distilled water must always be 20.00 cm^3 .

Do **not** carry out an experiment using 15.00 cm^3 of **FA 5**. Do **not** use a volume of **FA 5** that is less than 5.00 cm^3 .

Record all your results in a single table. You should include the volume of **FA 5**, the volume of distilled water and the reaction time.

The relative rate for the reaction is given by the following expression.

relative rate = $\frac{1000}{\text{reaction time in seconds}}$

Use this expression to calculate the relative rate for each of your experiments and record the values, to 3 significant figures, in your results table.

(i) Results

experiment	volume of FA 5 used / cm ³	volume of distilled water used / cm ³	reaction time / s	relative rate / s ⁻¹
1	20.00	0.0	17	58.8
2	10.00	10.0	35	28.6
3	5.00	15.0	68	14.7
4	7.50	12.5	40	25.0
5	12.50	7.5	28	35.7

- (ii) Plot a graph of relative rate against the volume of **FA 5** used on the grid in Fig 2.1. Include the origin in your plot. Draw a best–fit line. Label any points you consider anomalous. [3]
- (c) From your graph, what conclusion can you make about the relationship between the relative rate for the reaction and the volume of **FA 5** used? Explain your answer. [2]

Straight line passing through origin:

Rate is directly proportional to volume of **FA 5** used because straight line passes through the origin.

OR

Straight line not passing through origin:

Rate is proportional to volume of **FA 5** used as shown by the straight line drawn but not passing through origin.

OR

Curved line:

Rate is proportional to volume of **FA 5**. But it is not directly proportional because not a straight line (e.g. when **FA 5** used increases, rate increases at a decreasing rate).

(d) Explain why a small and fixed amount of sodium thiosulfate is required for this reaction. [2] The experimental rate (i.e. average rate) <u>approximates to the initial rate</u> of the reaction. AND

To ensure that the <u>same amount of iodine is formed</u> before the solution turns dark blue so that the all experiments <u>occur to the same extent</u> and rate of reaction is inversely proportional to time taken for solution to turn blue–black.

- (e) A student carried out the same experiment but used 15.00 cm³ of **FA5**. The student recorded a value for the reaction time of 28 s.
 - (i) Use your graph to determine the expected reaction time, to the nearest second, you would have to record if you had carried out an experiment using 15.00 cm³ of **FA 5**.

Show the construction lines on your graph and show your working in the calculation.

(read from graph) relative rate when 15.00 cm³ of **FA 5** was used = **A** s⁻¹ time taken = $\frac{1000}{A}$ = **B** s

(ii) Calculate the difference between your value and that of the student and express it as a percentage of the student's value. [1]

% difference = $\frac{28 - B}{28} \times 100\%$ (or $\frac{B - 28}{28} \times 100\%$) = **C** %

(f) (i) Kinetics studies have shown that the reaction in equation 1 is second order with respect to iodide ions and third order overall.

Steps 1 to 3 represent a possible mechanism of the reaction.

step 1	Fe ³⁺ + I [−]		[FeI] ²⁺
step 2	[FeI] ²⁺ + I [−]	\rightarrow	Fe ²⁺ + I ₂ ⁻
step 3	$2I_2^-$	\rightarrow	I₂ + 2I [−]

(i) State which step is the rate-determining step.

Step 2

(ii) The species in the steps shown have various roles. They can be reactants, products, catalysts or intermediates.

Suggest, with a reason, the role of [FeI]²⁺ ions in the mechanism shown above. [1]

 $[FeI]^{2+}$ ions are intermediate as they are <u>formed in step 1</u> and <u>used up / consumed in</u> <u>step 2</u> (and not regenerated in step 3).

[Total: 18]

[1]

[2]

3 Investigation of some inorganic reactions

At each stage of any test you are to record details of the following:

- colour changes seen;
- the formation of any precipitate and its solubility in an excess of the reagent added;
- the formation of any gas and its identification by a suitable test.

You should indicate clearly at what stage in a test a change occurs. If there is no observable change, write **no observable change**.

Unless otherwise stated, the volumes given below are approximate and should be estimated rather than measured. Rinse and reuse test-tubes, where possible.

No additional tests for ions present should be attempted.

- (a) FA 8 is a solution containing two cations. The two cations have different oxidation states.
 - (i) Carry out the following tests and record your observations in Table 3.1.

Table 3.1

test	observations			
To a portion of the solution FA 8 , add aqueous sodium hydroxide, until no further change is seen.	Pale blue ppt formed [✓], insoluble in excess NaOH [✓]			
Filter the mixture and add dilute sulfuric acid to the filtrate.	Colourless filtrate obtained. $[\checkmark]$ (pale blue residue obtained) white ppt reappears when dilute H ₂ SO ₄ is added, $[\checkmark]$ and the ppt is soluble in excess H ₂ SO ₄ to give a colourless solution $[\checkmark]$			
Wash the residue by pouring distilled water through it. Discard the washings and then pour dilute aqueous ammonia through the residue.	blue residue dissolved in excess NH₃ to give a dark blue solution [✓]			

[3]

(ii) The cations in **FA 8** are <u> Al^{3+} </u> and <u> Cu^{2+} </u>.

[2]

The deep blue solution obtained on adding excess $NH_3(aq)$ allows you to deduce the presence of Cu^{2+} . Both Zn^{2+} and Al^{3+} give the same observation recorded in Table 3.1 but Zn^{2+} has the same oxidation state as Cu^{2+} so Al^{3+} is the other cation.

(b) **FA 9** is a solution which contains two of the following anions: carbonate, chloride, bromide, iodide, sulfate(VI).

You are to introduce dilute nitric acid into **FA 9** prior to conducting further tests on the suspected anions present.

(i) To 2 cm depth of **FA 9** in a test-tube, add nitric acid.

Record your observations and deduce the identity of the anion present

Keep this solution for (b)(iii).

[1]

Effervescence produced. Gas evolved formed white ppt in limewater. $\mbox{CO}_3{}^{2-}$

(ii) By reference to the Quanlitative Analysis Notes, describe **two different tests**, using only the bench reagents provided, which will allow you to identify the other anion in **FA 9**.

test 1.....

st 2	test 2
[2]	

Test 1: add silver nitrate to a sample of **FA 9**, observe the colour of the precipitate. Add aq NH_3 to the precipitate.

If a white ppt forms and is soluble in NH_3 , C/ ions are present. If a pale cream ppt forms and is partially soluble in NH_3 , Br ions are present. If a yellow ppt forms and is insoluble in NH_3 , I^- ions are present.

Test 2: add barium nitrate to a sample of **FA 9**, observe if a precipitate forms. If a white ppt is formed, SO_4^{2-} ions are present.

(iii) Perform the tests you described in (b)(ii) on your solution from (b)(i).

Record your observations and hence deduce the identity of the other anion present.

Test 2: No ppt formed.

identity of anion: Chloride

4 Planning

When a monobasic acid, HX(aq) is mixed with aqueous potassium hydroxide, KOH(aq), the reaction releases heat causing a rise in the temperature of the solution.

$$HX(aq) + KOH(aq) \rightarrow KX(aq) + H_2O(I) \qquad \Delta H_1 < 0$$

A series of experiments can be performed where increasing volumes of HX(aq) and decreasing volumes of KOH(aq) are mixed and the temperature rise, ΔT , for each experiment is determined.

In each of the experiments using different volumes of HX(aq) and KOH(aq), the total volume has to be kept constant. Since the total volume of mixture remains the same, the temperature rise, ΔT , is a direct measure of the heat given out by the reaction. The maximum amount of heat is evolved when all the acid present is exactly neutralised by all the alkali present.

Plotting a graph of ΔT against the volume of HX used will give 2 straight lines of best–fit.

Extrapolation of the two straight lines will produce a point of intersection from which the concentration of HX(aq) and the enthalpy change, ΔH_1 , of the exothermic reaction between HX(aq) and KOH(aq) can be determined.

The volume of HX(aq) used should be at least 10.00 cm³ and the total volume of the reaction mixture should be kept constant at 50.00 cm³ for all experiments.

(a) Using the information given, you are required to write a plan to determine the concentration of HX and the enthalpy change, ΔH_1 , of the exothermic reaction between HX and sodium hydroxide.

You may assume that you are provided with:

- strong monobasic acid, HX, solution of unknown concentration,
- 2.0 mol dm⁻³ potassium hydroxide, KOH,
- Styrofoam cups and lids,
- the equipment normally found in a school or college laboratory.

In your plan you should include brief details of:

- the apparatus you would use,
- the quantities you would use (including suggested volumes of KOH and HX in a suitable table),
- the procedure you would follow,
- the measurements you would make,
- show how the results will be tabulated

[5]

Procedure

- 1 Place a Styrofoam cup inside a 250 cm³ glass beaker to prevent it from tipping over.
- 2 Use a <u>burette</u> to transfer 10.00 cm³ of **HX** into a <u>Styrofoam cup</u>.
- 3 Stir the solution in the cup with a <u>thermometer</u> and measure the <u>initial temperature of the</u> <u>HX</u> solution.
- 4 Rinse and wipe dry the thermometer.
- 5 Use a second <u>burette</u> to transfer 40.00 cm³ of KOH into a second Styrofoam cup.
- 6 Using the same thermometer, stir the solution gently and measure the <u>initial temperature</u> <u>of the **KOH**</u> solution.
- 7 Slip the thermometer through the lid of the cup.
- 8 Add **KOH** to **HX** in the cup and secure the lid onto the cup. Use the thermometer to stir the mixture and measure the maximum temperature of the mixture, T_{max} .
- 9 Wash and carefully dry both the Styrofoam cups used to contain the solutions.
- 10 <u>Repeat steps 1 to 8</u> above using appropriate volumes of **HX** and **KOH** as shown in Table 1 below so that the total volume of the reaction mixture is 50.00 cm³.
- 11 Record all measurements of volume, temperature and temperature change, ΔT , for the experiments, in Table 1 below.
- 12 For each experiment, calculate the <u>weighted average initial temperature</u>, T_{av} , of the two <u>solutions</u> used. The formula for T_{av} is given below:

$$T_{av} = \frac{\text{(volume HX \times initial temp HX)} + \text{(volume KOH \times initial temp KOH)}}{50.00}$$

Table 1							
No	V _{HX}	V KOH	initial	initial	T () C	T () C	
INO.	/ cm ³	/ cm ³	of HX / °C	of KOH / °C	av / C	max / °C	
1	10.00	40.00					
2	15.00	35.00					
3	20.00	30.00					
4	25.00	25.00					
5	30.00	20.00					
6	35.00	15.00					
7	40.00	10.00					

$$\Delta T = T_{\max} - T_{av}$$

- (b) Sketch on Fig 4.1 the graph you would expect to obtain from your results. Indicate clearly on your sketch how you would determine:
 - *V*_{neut}, the volume of HX(aq) needed to just completely neutralise (50 *V*_{neut}) cm³ of KOH(aq).
 - Δ*T*_{max}, the maximum temperature rise when stoichiometric amount of HX(aq) and KOH(aq) reacted.

[2]

[1]



- (c) Outline how you would use your answers from (b) to determine
 - (i) the concentration, in mol dm^{-3} , of HX(aq).

You may show your results in terms of V_{neut}

$$n(HX) = n(KOH)$$
Vol of HX = V_{neut} cm³ Vol of NaOH = (50 - V_{neut}) cm³
Amt of NaOH = Amt of HX = $(\frac{50 - V_{neut}}{1000}) \times 2$ OR $\frac{100 - 2V_{neut}}{1000}$ OR 0.1 - 0.002V_{neut}
[HX] = $\frac{\frac{50 - V_{neut}}{1000}}{\frac{V_{neut}}{1000}}$ mol dm⁻³ OR $\frac{100 - 2V_{neut}}{V_{neut}}$ OR $\frac{100}{V_{neut}} - 2$

(ii) the enthalpy change of reaction, ΔH_1 , for this reaction.

$$HX(aq) + KOH(aq) \rightarrow KX(aq) + H_2O(l)$$
 ΔH_1

Express your answers in terms of ΔT_{max}

You should assume that the specific heat capacity of the final solution is $4.18 \text{ J g}^{-1} \text{ K}^{-1}$ and its density is 1.00 g cm^{-3} .

Heat change of reaction = (50 cm³ x 1.00 g cm⁻³ x 4.18 J g⁻¹ K⁻¹ x ΔT_{max} °C) J = 50 x 4.18 x ΔT_{max} = 209 ΔT_{max} J

(Correct calculation of heat change using ΔT_{max} . Ignore sign if not given. Allow kJ.) From (c)(i) amt of HX or KOH = $(\frac{50 - V_{neut}}{1000}) \times 2$ OR $\frac{100 - 2V_{neut}}{1000}$ OR $0.1 - 0.002V_{neut}$ $\Delta H_1 = -\frac{\text{heat change of reaction}}{n(\text{HX})}$ kJ mol⁻¹ Must have negative sign.

(d) A similar exothermic reaction will take place between a monobasic **weak** acid, HY(aq) and KOH(aq).

Explain how each of the following

- V_{neut}(HY), the volume of HY(aq) needed to just completely neutralise (50 - V_{neut}(HY)) cm³ of KOH(aq).
- $\Delta T_{max}(HY)$, the maximum temperature rise when stoichiometric amount of HY(aq) and KOH(aq) reacted.

would differ from the values obtained in (b) when the same experiment you have described in (a) is repeated using a monobasic weak acid, HY(aq) of the same concentration.

[2]

[2]

- Same V_{eq}. HY(aq)
 H⁺(aq) + Y⁻(aq). When KOH is added, it will neutralise H⁺ and thus [H⁺] will decrease and equilibrium position will shift to the right allowing more H⁺ to dissociate. Thus, all weak acid HY will react with the KOH resulting in the same volume of acid used.
- Lower △*T*_{max}. Since HY is a weak acid, more energy will be absorbed to allow complete dissociation of H⁺

[Total: 14]