

DUNMAN HIGH SCHOOL Preliminary Examination Year 6

H2 CHEMISTRY

Paper 4 Practical

9729/04 24 August 2023 2 hours 30 minutes

Candidates answer on the Question Paper.

READ THESE INSTRUCTIONS FIRST

Write your centre number, index number, name and class at the top of this page.

Give details of the practical shift and laboratory where appropriate, in the boxes provided. Write in dark blue or black pen.

You may use an HB pencil for any diagrams or graphs.

Do not use staples, paper clips, glue or correction fluid.

Answer all questions in the spaces provided on the Question Paper.

The use of an approved scientific calculator is expected, where appropriate.

You may lose marks if you do not show your working or if you do not use appropriate units.

Qualitative Analysis Notes are printed on pages 19 and 20.

At the end of the examination, fasten all your work securely together. The number of marks is given in brackets [] at the end of each question or part question.

Shift
Laboratory

For Examiner's Use		
1		
	16	
2		
	13	
3		
_	15	
4		
	11	
Total		
	55	

2

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

1 Determination of change in oxidation number in XO_3^- oxidant.

FA 1 is 0.0265 mol dm⁻³ potassium manganate(VII), KMnO₄(aq) **FA 2** is 0.0750 mol dm⁻³ XO_3^- oxidant solution. X is the symbol of an element whose identity is not needed in this question. **FA 3** is a solution containing 0.150 mol dm⁻³ iron(II) ions, Fe²⁺.

FA 3 solution is prepared by dissolving an iron(II) salt in sulfuric acid.

In the presence of acid, XO_3^- oxidises iron(II) to iron(III) while itself is reduced to X^{b+} . In order to determine *b*, a sample of XO_3^- is added to a known amount of iron(II) in excess. The amount of remaining iron(II) ions can then be determined by titration against manganate(VII) ions as shown in equation 1.

equation 1 $MnO_4^{-}(aq) + 8H^{+}(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(I)$

(a) Procedure

- 1. Pipette 25.0 cm³ of **FA 3** into a conical flask.
- 2. Using another pipette, add exactly 10.0 cm³ of **FA 2** to **FA 3** in the conical flask.
- 3. Fill the burette with **FA 1**.
- 4. Titrate the mixture in the conical flask with **FA 1** until the appearance of the first permanent pale orange or pale pink colour.
- 5. Repeat the titration as many times as you think necessary to obtain accurate results.
- 6. Record your results of titration in a suitable tabulated form below.

Results

(b) From your titrations, obtain a suitable volume of **FA 1**, *V*_{FA 1}, to be used in your calculations. Show clearly how you have obtained this volume.

V_{FA 1} = cm³ [3]

(c) (i) Calculate the amount of Fe^{2+} ions pipetted into the conical flask.

(ii) Calculate the amount of XO_3^- ions present in the conical flask.

amount of XO_3^- ions = [1]

(iii) Using your answer to (b), calculate the amount of Fe^{2+} ions that remain after reaction with XO_3^{-} .

(iv) Calculate the amount of Fe^{2+} ions that reacted with XO_3^- ions in the conical flask and hence, the amount of Fe^{2+} that reacted with one mole of XO_3^- ions.

(v) Hence, calculate the final oxidation number of the element X when Fe²⁺ reacts with XO_3^- .

final oxidation number of the element $X = \dots$ [2]

(d) A student carries out the same experiment in (a). However unknowingly, he uses a sample of FA 2 with concentration 0.100 mol dm⁻³ instead.

Suggest what effect this will have on the volume of FA 1 obtained. Explain your answer.

.....[1]

(e) (i) A student followed the procedure described in (a) using a different oxidant instead of FA 2. The student obtained a mean titre value of 22.20 cm³. The actual volume of FA 1 required should have been 22.40 cm³.

Calculate the difference between the actual value and the student's value as a percentage. This is the experimental error.

experimental error =% [1]

(ii) Using his mean titre value, the student calculates that the maximum total percentage error of the apparatus used is 0.77%.

Use the following data to verify the student's calculation.

apparatus	uncertainty
25 cm ³ pipette	±0.03 cm ³
10 cm ³ pipette	±0.02 cm ³

Hence, determine if the student performed the experiment well. Explain your answer.

[Total: 16]

2 Determination of the kinetics of the reaction between Fe³⁺ ions and iodide ions, I⁻.

You are provided with the following reagents.

FA 4 contains 0.0250 mol dm⁻³ acidified iron(III) chloride, FeC l_3 **FA 5** is 0.0800 mol dm⁻³ aqueous potassium iodide, KI **FA 6** is 0.00500 mol dm⁻³ sodium thiosulfate, Na₂S₂O₃ starch solution

Fe³⁺ ions oxidise iodide ions, I⁻, to iodine, I₂, as shown in equation 2. In this experiment, you will investigate how the rate of this reaction is affected by the concentration of Fe³⁺ ions.

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

A fixed and small amount of thiosulfate ions, $S_2O_3^{2-}$, and starch indicator will be added to a mixture of Fe³⁺(aq) and I⁻(aq). The iodine, I₂, produced reacts immediately with thiosulfate ions, $S_2O_3^{2-}$, as shown in equation 3.

equation 3 $I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^{-}(aq) + S_4O_6^{2-}(aq)$

When all the thiosulfate has been used, the iodine produced will turn starch indicator blue-black. The rate of the reaction can therefore be measured by finding the time it takes for the reaction mixture to turn blue–black.

You will perform a series of **four** experiments. In the series of experiments, the rate equation for the reaction can be simplified to rate = k'[Fe³⁺]^{*n*}, where *n* is the order of reaction with respect to [Fe³⁺] and *k'* is *k*[I⁻]². You will graphically analyse your results to determine *n*.

For each experiment, you will note the volume of **FA 4** added, $V_{FA 4}$, and the time taken, *t*, to the nearest second, for the reaction mixture to become blue-black. In each experiment, you will need to ensure that the same total volume of reaction mixture is used by adding deionised water as required.

- (a) Prepare a table in the space provided on page 8 in which to record, to an appropriate level of precision:
 - volumes of FA 4, V_{FA 4}, and deionised water,
 - all values of *t*,

• all calculated values of
$$\frac{1}{t}$$
, $Ig\left(\frac{1}{t}\right)$ and $Ig(V_{FA 4})$.

(i) Experiment 1

- 1. Fill a burette with **FA 4**.
- 2. Transfer 20.00 cm³ of **FA 4** into a 100 cm³ conical flask.
- 3. Use appropriate measuring cylinders, place the following in a 100 cm³ beaker.
 - 10.0 cm³ of **FA 5**
 - 15.0 cm³ of **FA 6**
 - 10.0 cm³ of starch solution
- 4. Add the contents of the beaker rapidly to the conical flask and start the stopwatch.
- 5. Swirl the mixture and place the conical flask on the white tile.
- 6. The mixture turns purple and then yellow before turning a blue-black colour. Stop timing when this **blue-black colour** appears.
- 7. Record the time taken, *t*, to the nearest second in your table.
- 8. Discard the reaction mixture **immediately** down the sink. Wash out the conical flask and stand it upside down on a paper towel to drain.

(ii) Experiment 2

- 1. Run 6.00 cm³ of **FA 4** into a 100 cm³ conical flask.
- 2. Using another burette, add 14.00 cm³ of deionised water into the conical flask containing **FA 4**.
- 3. Use the measuring cylinders from **experiment 1** to place the following in a 100 cm³ beaker.
 - 10.0 cm³ of **FA 5**
 - 15.0 cm³ of **FA 6**
 - 10.0 cm³ of starch solution
- 4. Add the contents of the beaker rapidly to the conical flask and start the stopwatch.
- 5. Swirl the mixture and place the conical flask on the white tile.
- The mixture turns purple and then yellow before turning a blue-black colour. Stop timing when this blue-black colour of similar intensity as experiment 1 appears.
- 7. Record the time taken, *t*, to nearest second in your table.
- 8. Discard the reaction mixture **immediately** down the sink. Wash out the conical flask and stand it upside down on a paper towel to drain.

In **experiment 1** you will have obtained the time taken for a 'fast' reaction and in **experiment 2** the time taken for a 'slow' reaction.

Carry out **two** further experiments to investigate the effect of changing the concentration of $Fe^{3+}(aq)$ by altering the volume of Fe^{3+} , **FA 4**, used. In each case, you will need to ensure that the **same total volume** of reaction mixture is used by adding deionised water as required.

You should alternate the use of the two 100 cm³ conical flasks.

Record all required volumes, time taken and calculated values in your table.

(iii) Results

(b) (i) Plot a graph of $Ig\left(\frac{1}{t}\right)$ on the y-axis against $Ig(V_{FA 4})$ on the x-axis. Draw the best-fit straight line taking into account all of your plotted points.

(ii) Calculate the gradient of the line to three significant figures, showing clearly how you did this. Hence deduce n, the order of the reaction with respect to $[Fe^{3+}]^n$ in rate = $k'[Fe^{3+}]^n$.

9729/04

order of the reaction with respect to $[Fe^{3+}] = \dots$ [3]

(c) (i) A student wants to investigate the effect of concentration of iodide ions on the rate of this reaction. He plans to carry out another one experiment, **experiment 5.**

Suggest appropriate volumes of chemicals to be used in the procedure by completing Table 2.1 so that his results in **experiment 5**, together with that from **experiment 1**, could be used to determine the order of reaction with respect to $[I^-]$.

experiment	volume of FA 4 / cm ³	volume of FA 5 / cm ³	volume of FA 6 / cm ³	volume of deionised water / cm ³	time taken for blue-black to appear / s
1	20.00	10.0	15.0	0.0	t ₁
5					t ₂

Table 2.1

(ii) Taking the time for the blue-black colour to appear in **experiments 1 and 5** to be t_1 and t_2 respectively, explain how the results of the two experiments can be used to confirm that this reaction is second order with respect to $[I^-]$.

[1]

(iii) Suggest an improvement to the experimental procedure so that the order of reaction with respect to [I⁻] can be more **reliably** determined.

[1] [Total: 13]

3 Qualitative Analysis

(a) FA 7 contains two cations.
FA 5 is 0.0800 mol dm⁻³ aqueous potassium iodide, KI, provided in Question 2.
FA 8 is a solution of a dilute acid.

11

Perform the tests described in Table 3.1 and record your observations in the spaces provided.

Your answers should include:

- details of colour changes and precipitates formed.
- the names of gases evolved and details of the test used to identify each gas.

Table 3	3.1
---------	-----

	tests	observations
(i)	To 1 cm depth of FA 7 in a test– tube, add an equal depth of aqueous ammonia.	
(ii)	To 1 cm depth of FA 7 in a test- tube, add one spatula of zinc powder. Stir well. Leave to stand for a few minutes.	
(iii)	To 1 cm depth of FA 7 in a test– tube, add about 2 cm depth of FA 5 and shake the mixture thoroughly. Then add a few drops of starch.	
(iv)	To 1 cm depth of FA 7 in a boiling tube, add aqueous sodium hydroxide.	
	Warm the mixture.	

[5]

(b) Identify the cations present in **FA 7**. Use evidence from your observations in (a) to support your deduction.

cation	evidence

[2]

(c) With the aid of chemical equations, explain your observations for the test in (a)(i).

 	 	[3]

(d) Using your observations in test (a)(iii), state and explain the role of FA 7.

.....[1]

- (e) The identity of FA 8 could be hydrochloric acid, nitric acid or sulfuric acid.
 - (i) Devise a series of tests that will identify **FA 8**. For each of the possible acids, you should indicate the expected observations for a **positive test** in Table 3.2. Your tests should be based on the Qualitative Analysis Notes on pages 19–20 and should use only the bench reagents provided.

	test	expected observations for a positive test
1		
2		
3		

Table 3.2

[3]

(ii) Perform your tests on **FA 8** to confirm its identity.

Any test requiring heating MUST be performed in a boiling tube.

FA 8 is

[1]

[Total: 15]

4 Planning: Determining the equilibrium constant, *K*_c, for a reaction

The hydrolysis of an ester is catalysed by either an acid or alkali. In the presence of an acid catalyst, the reaction is reversible and reaches equilibrium in three days.

Ethyl ethanoate hydrolyses to give ethanoic acid and ethanol as shown in equation 4.

equation 4 $CH_3CO_2CH_2CH_3(I) + H_2O(I) \implies CH_3CO_2H(I) + CH_3CH_2OH(I)$

To determine the equilibrium constant, K_c , it is necessary to determine the amounts of ethyl ethanoate, water, ethanoic acid and ethanol present in the equilibrium mixture. One way of doing this is described below.

Known amounts of ethyl ethanoate and an acid catalyst are mixed. After sufficient time, about 50 cm³ of deionised water is added to the equilibrium mixture. The whole mixture is then titrated against a standard solution of sodium hydroxide. This allows the total amount of acid in the mixture to be found. The total amount of acid is made up of the acid catalyst and the ethanoic acid present in the equilibrium mixture.

Since the amount of acid catalyst does not change, it is possible to deduce the amount of ethanoic acid in the equilibrium mixture. If the initial amounts of ethyl ethanoate and water are known, the amounts of all species present in the equilibrium mixture can be determined.

These data could then be used in the following expression to determine K_c .

 $\mathcal{K}_{c} = \frac{[ethanoic acid][ethanol]}{[ethyl ethanoate][water]}$

(a) Using the information given above, write a plan to determine the equilibrium constant, K_c , for the hydrolysis of ethyl ethanoate at 35 °C.

In your plan, you should make an initial reaction mixture containing

- 5.00 cm³ of ethyl ethanoate,
- 5.00 cm³ of 3.00 mol dm⁻³ hydrochloric acid.

You are also provided with:

- 1.00 mol dm⁻³ sodium hydroxide,
- thymolphthalein indicator (colourless in acidic medium, blue in alkaline),
- the equipment and materials normally found in a school or college laboratory.

You should include brief details of:

- the apparatus you would use,
- the procedure you would use to prepare the equilibrium mixture,
- the procedure you would use to determine the total amount of acid in the equilibrium mixture,
- the measurements you would make.

In your plan, you will titrate the **whole** equilibrium mixture so only **one** titration can be carried out. You do not need to describe any calculations.

.....[6] (b) Table 4.1 gives information about the densities of the reactants.

name	Mr	density / g cm ⁻³
water	18.0	1.00
ethyl ethanoate	88.0	0.90

|--|

Assume that 40.00 cm³ of 1.00 mol dm⁻³ sodium hydroxide was needed to neutralise the acids present in the equilibrium mixture, calculate the amounts (in moles) of

- ester and water present in your initial reaction mixture,
- ethanoic acid present in this mixture at equilibrium,
- all the other components in this mixture at equilibrium.

You may assume that 1.00 cm³ of hydrochloric acid contains 1.00 cm³ of water.

Hence, determine the value of equilibrium constant, K_c , for equation 4.

(c) Two other experiments were carried out using different volumes of ethyl ethanoate as shown in Table 4.2.

experiment	volume of ethyl ethanoate / cm ³	volume of hydrochloric acid / cm³	volume of water / cm ³
1	5.00	5.00	0.00
2	4.00	5.00	1.00
3	3.00	5.00	2.00

Sketch a graph in Fig. 4.1 to show how the value of K_c obtained for experiments 2 and 3 would compare with experiment 1.





[1]

[Total: 11]

BLANK PAGE

18

Qualitative Analysis Notes [ppt. = precipitate]

(a) Reactions of aqueous cations

	reaction with		
calion	NaOH(aq)	NH₃(aq)	
aluminium, Al ³⁺ (aq)	white ppt. soluble in excess	white ppt. insoluble in excess	
ammonium, NH₄⁺ (aq)	ammonia produced on heating	-	
barium, Ba ²⁺ (aq)	no ppt. (if reagents are pure)	no ppt.	
calcium, Ca ²⁺ (aq)	white. ppt. with high [Ca ²⁺ (aq)]	no ppt.	
chromium(III), Cr³+(aq)	grey–green ppt. soluble in excess giving dark green solution	grey–green ppt. insoluble in excess	
copper(II), Cu²+(aq)	pale blue ppt. insoluble in excess	blue ppt. soluble in excess giving dark blue solution	
iron(II), Fe ²⁺ (aq)	green ppt., turning brown on contact with air insoluble in excess	green ppt., turning brown on contact with air insoluble in excess	
iron(III), Fe ³⁺ (aq)	red–brown ppt. insoluble in excess	red–brown ppt. insoluble in excess	
magnesium, Mg ²⁺ (aq)	white ppt. insoluble in excess	white ppt. insoluble in excess	
manganese(II), Mn²+(aq)	off–white ppt., rapidly turning brown on contact with air insoluble in excess	off–white ppt., rapidly turning brown on contact with air insoluble in excess	
zinc, Zn²+(aq)	white ppt. soluble in excess	white ppt. soluble in excess	

(b) Reactions of anions

ion	reaction		
carbonate, CO ₃ ²⁻	CO ₂ liberated by dilute acids		
chloride, C <i>l</i> ⁻(aq)	gives white ppt. with Ag⁺(aq) (soluble in NH₃(aq))		
bromide, Br⁻(aq)	gives pale cream ppt. with Ag ⁺ (aq) (partially soluble in NH ₃ (aq))		
iodide, I⁻(aq)	gives yellow ppt. with Ag ⁺ (aq) (insoluble in $NH_3(aq)$)		
nitrate, NO₃⁻ (aq)	NH ₃ liberated on heating with OH ⁻ (aq) and A l foil		
nitrite	NH_3 liberated on heating with OH ⁻ (aq) and Al foil;		
NO ₂ ⁻ (aq)	NO liberated by dilute acids		
	(colourless NO \rightarrow (pale) brown NO ₂ in air)		
sulfate, SO₄²- (aq)	gives white ppt. with Ba ²⁺ (aq) (insoluble in excess dilute strong acids)		
sulfite,	SO ₂ liberated with dilute acids;		
SO ₃ ²⁻ (aq)	gives white ppt. with Ba ²⁺ (aq) (soluble in dilute strong acids)		

(c) Tests for gases

gas	test and test result	
ammonia, NH ₃	turns damp red litmus paper blue	
carbon dioxide, CO ₂	de, CO ₂ gives a white ppt. with limewater (ppt. dissolves with excess CO ₂)	
chlorine, Cl ₂	bleaches damp litmus paper	
hydrogen, H ₂	"pops" with a lighted splint	
oxygen, O ₂	relights a glowing splint	
sulfur dioxide, SO ₂	fur dioxide, SO ₂ turns aqueous acidified potassium manganate(VII) from purple to colourless	

(d) Colour of halogens

halogen	colour of element	colour in aqueous solution	colour in hexane
chlorine, Cl_2	greenish yellow gas	pale yellow	pale yellow
bromine, Br ₂	reddish brown gas / liquid	orange	orange-red
iodine, I2	black solid / purple gas	brown	purple