NATIONAL JUNIOR COLLEGE
SH 2 Year − End Practical Examination
Higher 2

CANDIDATE NAME			
SUBJECT CLASS		REGISTRATION NUMBER	

CHEMISTRY 9729/04

Paper 4 Practical **Wednesday 19 August 2020**

Candidate answer on the Question paper. **2 hours 30 minutes**

READ THESE INSTRUCTIONS FIRST

Write your identification number and name.
Give details of the practical shift and laboratory where appropriate, in the boxes provided.
Write in 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.

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	/ 13
4	/ 12
Presentation	/ 1
Total	/ 55

This paper consists of **18** printed pages including this cover page.

1 Determination of a value for the number of water of crystallization, *x*, in an iron(II) salt, $(NH_4)_2Fe(SO_4)_2.\ x\ H_2O$.

The iron (II) salt, **$(NH_4)_2Fe(SO_4)_2.\ x\ H_2O$**, can be titrated against potassium manganate (VII) solution. The two half–equations for the reaction are shown below.

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

$$Fe^{2+} \longrightarrow Fe^{3+} + e^-$$

FA 1 is potassium manganate(VII), $KMnO_4$ solution, of concentration **around 0.02 mol dm^{-3}**
FA 2 is 0.0780 mol dm^{-3} ethanedioate ions $C_2O_4^{2-}$
FA 3 is a solution of the iron(II) salt containing 38.4 g dm^{-3} of $(NH_4)_2Fe(SO_4)_2.\ x\ H_2O$
FA 4 is dilute sulfuric acid, H_2SO_4

The exact concentration of **FA 1** can be determined when **FA 2** is titrated against it.

$C_2O_4^{2-}$ reacts with MnO_4^- in an acidic medium to form Mn^{2+} and CO_2 gas.

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

As this reaction is very slow at the start, the $C_2O_4^{2-}$ solution is heated prior to titration against MnO_4^-. However, it is not necessary to continuously heat the solution as the rate of reaction will increase as the reaction proceeds.

Method

(a) (i) Standardisation of FA 1 solution

1. Fill the burette with **FA 1**.
2. Using a pipette, transfer 10.0 cm^3 of **FA 2** into the conical flask.
3. Using an appropriate measuring cylinder, transfer 20 cm^3 of **FA 4** to the same conical flask.
4. Place a 1 ^{o}C thermometer into this conical flask.
5. Heat the contents of the conical flask using the bunsen burner.
6. **Turn off the Bunsen burner** once a temperature of 60 ^{o}C is reached.

Caution: The conical flask should be comfortable to the touch. However, you may also use the paper towels to handle the neck of the heated flask if it is too hot for you.

7. Using a small volume of deionised water, wash down the sides of the thermometer and remove it from the conical flask.
8. Add **FA 1** from the burette into this flask until a **permanent** pale pink colour is obtained.

Initially, the colour of **FA 1** will take some time to disappear, continue swirling. Subsequently, as the reaction proceeds, the rate of reaction will increase.

9. Record your titration results in the space provided on page **3**.
10. Repeat steps **1** to **7** as necessary until consistent results are obtained.

Results:

	1	2
Final burette reading / cm^3	**15.10**	**29.90**
Initial burette reading / cm^3	**0.30**	**15.10**
Volume of FA 1 / cm^3	**14.80**	**14.80**

[1]: Both titration tables with correct headings and units

(Do not award if any final and initial burette readings are inverted / if 50.00 is used as initial burette reading / burette reading is > 50.00)

[1]: Both titration tables with all burette readings recorded to the nearest 0.05 cm^3 + correct computation of titres

[1]: Both tables with two uncorrected titres for endpoint within 0.10 cm^3.

[3]

(ii) From your titration results, obtain a suitable volume of **FA 1** to be used in your calculations. Show clearly how you obtained this volume.

Average volume of FA 1 used = (14.80 + 14.80) / 2 = 14.80 cm^3 (2 d.p.)

- Quote the selected titres in the working for average volume
- Give correct final value to 2 decimal places
- No need write statement, can just show working

Volume of **FA 1** used = **14.80 cm^3**

[1]

(iii) Hence, determine the concentration of MnO_4^- in **FA 1**.

Amount of $C_2O_4^{2-}$ used in 10.0 cm^3 = $0.0780 \times \frac{10.0}{1000}$ = 7.80×10^{-4} mol

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

$2MnO_4^- \equiv 5C_2O_4^{2-}$

Amount of MnO_4^- reacted in 14.80 cm^3 = $\frac{2}{5} \times 7.80 \times 10^{-4}$ = 3.12×10^{-4} mol **[1]**

$[MnO_4^-] = 3.12 \times 10^{-4} / \frac{14.80}{1000}$ = 0.0211 mol dm^{-3} (3 s.f.) **[1] (ecf from 1st point)**

Concentration of MnO_4^- in **FA 1** = **0.0211 mol dm^{-3}**

[2]

(iv) For the titration in **(a)(i)**, the $C_2O_4^{2-}$ solution needs to be at about 60 °C at the start.

During the titration, as **FA 1** is added, the temperature of the mixture decreases. However, this decrease in temperature did not cause the rate of reaction between $C_2O_4^{2-}$ and MnO_4^- to decrease.

(i) Suggest a reason for the slow reaction at the start.

The rate of reaction between MnO_4^- and $C_2O_4^{2-}$ is slow initially due to the **electrostatic repulsion between the two negatively charged ions,** resulting in high activation energy.

[1]

(ii) Explain why the rate of reaction did not decrease even though the temperature of the mixture decreased.

As the reaction proceeds, **Mn^{2+}, a catalyst is generated** which increases the rate of the reaction OR **autocatalysis reaction** has taken place.

[1]

(v) Calculate the maximum total percentage uncertainty for your titration in **(a)(i)** if the uncertainty associated with each reading using a 10.0 cm^3 pipette and a burette are ±0.03 cm^3 and ±0.05 cm^3 respectively.

% uncertainty due to pipette = $\frac{\pm 0.03}{10.0}$ x 100 % = ± 0.300 %

% uncertainty due to burette used in titration = $\frac{\pm 0.05 \times 2}{14.80}$ x 100 % = ± 0.676 %

Maximum total percentage uncertainty = 0.300 + 0.676 = ± 0.976 % (3 s.f.)

(0 mark if ± missing)

Percentage uncertainty = **± 0.976 %**

[1]

(b) (i) Titration of FA3 against FA1

1. Fill the burette with **FA 1**.
2. Pipette 25.0 cm^3 of **FA3** into a conical flask and add 20 cm^3 of **FA 4** using an appropriate measuring cylinder.
3. Titrate **FA 3** with **FA 1** until a pale orange colour is obtained.
4. Carry out as many titrations as you think necessary to obtain consistent results. Record your titration results in the space on the next page.

Results:

	1	2
Final burette reading / cm^3	26.55	35.90
Initial burette reading / cm^3	2.90	12.20
Volume of FA 1 / cm^3	23.65	23.70

Accuracy: ± 0.20 cm^3 of teacher's reading: **[2]**
± 0.40 cm^3 of teacher's reading: **[1]**

[2]

(ii) From your titration results, obtain a suitable volume of **FA 1** to be used in your calculations. Show clearly how you obtained this volume.

Average volume of FA 1 used = (23.65 + 23.70) / 2 = 23.68 cm^3 (2 d.p.)

- Quote the selected titres in the working for average volume
- Give correct final value to 2 decimal places
- No need write statement, can just show working

Volume of **FA 1** used = **23.68 cm^3**

[1]

(c) Calculations

(i) Calculate the amount of iron(II) ions present in 1.00 dm^3 of **FA 3**.

Amount of $KMnO_4$ in 23.68 cm^3 = $\frac{23.68}{1000} \times 0.02108 = 4.992 \times 10^{-4}$ mol

$MnO_4^- + 8H^+ + 5Fe^{2+} \longrightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$

$5Fe^{2+} \equiv MnO_4^-$

Amount of Fe^{2+} in 25.0 $cm^3 = 5 \times 4.992 \times 10^{-4} = 2.496 \times 10^{-3}$ mol **[1]**

Amount of Fe^{2+} in 1.00 $dm^3 = 2.496 \times 10^{-3} \times \frac{1000}{25.0}$ = 0.0998 mol (3 s.f.)

[1] (ecf from 1st point)

Amount of iron(II) ions in 1.00 dm^3 of **FA 3** = **0.0998 mol**

[2]

(ii) Using your answer in **(c)(i)**, determine the value of x, the number of water of crystallisation in the iron(II) salt, $(NH_4)_2Fe(SO_4)_2.\ x\ H_2O$.

[A_r: Fe, 55.8; H, 1.0; N, 14.0; O, 16.0; S, 32.1]

Fe^{2+} ≡ iron(II) salt ➔ Amount of iron(II) salt in 1.00 dm^3 of FA 3 = 0.09983 mol

Formula mass of iron(II) salt = $\frac{38.4}{0.09983}$ = 384.7 g mol^{-1} **[1]**

Relative formula mass of iron(II) salt = $2(14.0 + 4.0) + 55.8 + 2(32.1 + 64.0) + x(2.0 + 16.0)$

$384.7 = 284.0 + 18x$

$100.7 = 18x$

$x = 5.59 \approx 6$ (to nearest integer) **[1] (ecf from 1st point)**

x = **6**

[2]

[Total: 16]

2 Determination of a value for the enthalpy change, ΔH, of the reaction between magnesium and hydrochloric acid.

Magnesium reacts with an excess of hydrochloric acid according to the following equation.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

FA 5 is 2 mol dm^{-3} hydrochloric acid, HC*l*

FA 6 is magnesium ribbon, Mg

You are also provided with sandpaper

(a) Method

1. Place a Styrofoam cup in a 250 cm^3 beaker to prevent it from tipping over.

2. Sand a piece of **FA6**, weigh and record its mass in the space below.

3. Using a measuring cylinder, transfer 25 cm^3 of **FA 5** into the Styrofoam cup.

4. Measure the temperature of **FA 5** in the Styrofoam cup. This is the temperature at time = 0 minute.

5. Start the stopwatch and leave it running for the whole experiment.

6. Record the temperature of the acid every half a minute for 2 minutes.

7. At time = 2.5 minutes, carefully drop the piece of **FA 6** into the acid and stir the mixture. Do not measure the temperature.

8. Record the temperature of the mixture at time = 3 minutes, then every half a minute over a total duration of 8 minutes. Stir the mixture between thermometer readings.

Results:

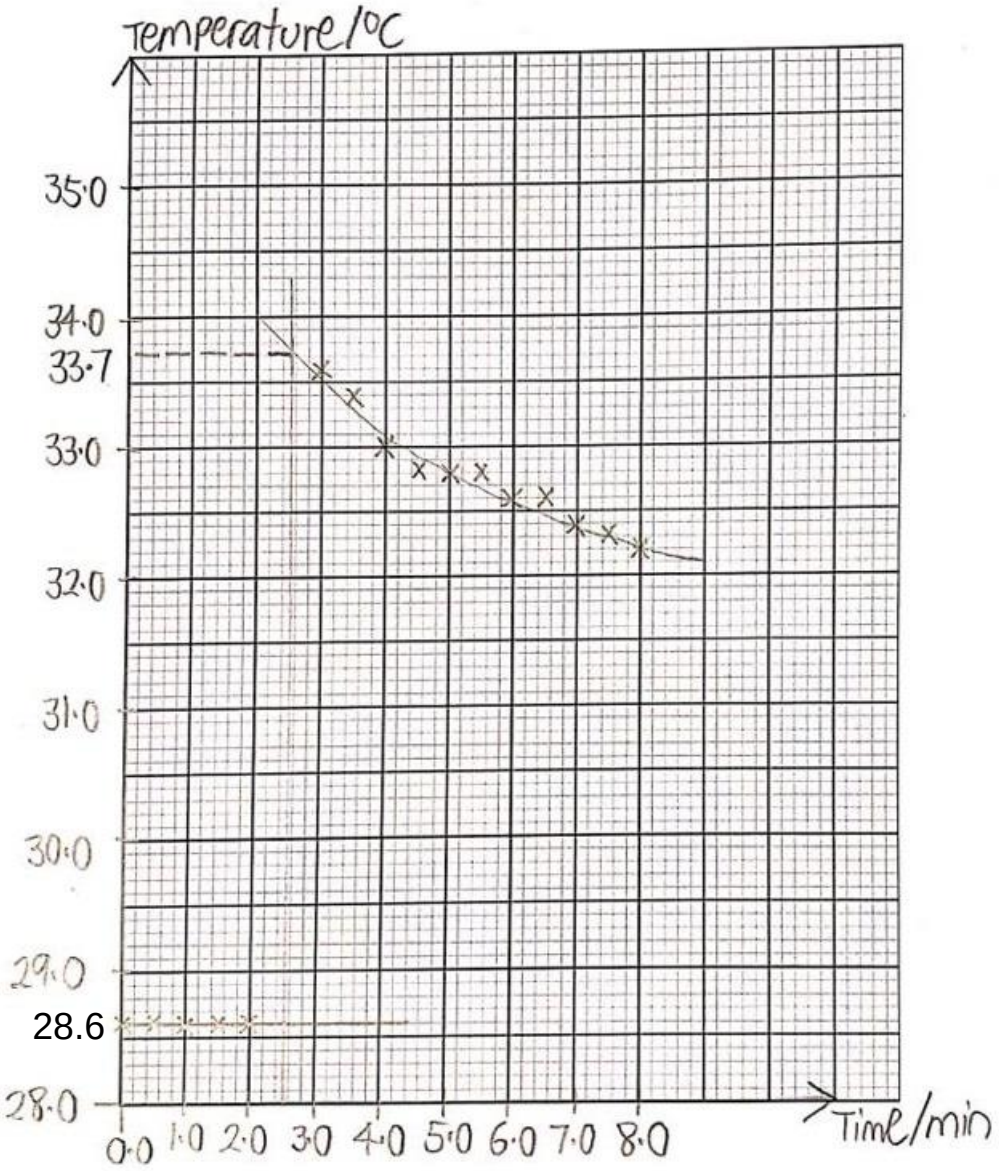
Mass of Mg ribbon	0.032 g **[1]**

Time / min	Temp. / °C	Time / min	Temp. / °C
0.0	28.6	4.5	32.8
0.5	28.6	5.0	32.8
1.0	28.6	5.5	32.8
1.5	28.6	6.0	32.6
2.0	28.6	6.5	32.6
2.5	–	7.0	32.4
3.0	33.6	7.5	32.3
3.5	33.4	8.0	32.2
4.0	33.0		

[1] correct headings with units
[1] no temperature reading at 2.5 min
[1] complete data with 1 decimal place for temperature values

[4]

(b) Plot a graph of temperature against time on the grid below.



Points plotted within half a square of 'x' **[1]**

Curve and line of best fit, points are within 1.5 squares from line **[1]**

Axis calibrated and labelled correctly **[1]**

Deduct 1 mark if graph covers less than 50% of each axis

[3]

Draw two lines of best fit through the points on your graph, the first for the temperature before adding **FA 6** and the second for the cooling of the mixture once the reaction is complete.

Extrapolate the two lines to time = 2.5 minutes, draw a vertical line between the two and determine the theoretical rise in temperature at this time.

Theoretical rise in temperature at 2.5 minutes = **5.1** °C [1]

(c) Calculation

(i) Use your answer to **(b)** to calculate the heat energy, in joules, given out when **FA 6** is added to **FA 5**.

(Assume 4.2 J of heat energy raises the temperature of 1.0 cm^3 of the mixture by 1.0 °C)

$\Delta q_{surroundings} = mc\Delta T = 25 \times 4.2 \times 5.1 = 535.5\ J = 536\ J$ (3 s.f.)

Heat energy evolved = **536** J

[1]

(ii) Use your answer to **(i)** to calculate the enthalpy change, in kJ mol^{-1}, when 1 mole of **FA 6** reacts with **FA 5**.

[A_r: Mg, 24.3; H, 1.0; C*l*, 35.5]

Amount of Mg = 0.032 / 24.3 = 0.001317 mol **[1]**

ΔH = −535.5 / 0.001317 = −407 kJ mol^{-1} **[1] (ecf from 1st point)**

0 mark if no negative sign

Enthalpy change = **−407** kJ mol^{-1}

[2]

(d) A student carried out the same procedure using the same concentration of sulfuric acid, H_2SO_4, instead of hydrochloric acid and predicted that the enthalpy change would be twice of that with hydrochloric acid. Comment on this prediction.

The student is incorrect as the enthalpy change of this reaction is defined by the **amount of heat released per mole of magnesium** (as seen in part **cii**).

Enthalpy change is not affected by the individual amounts of the reactants.

[1]

(e) The enthalpy change determined in **(ii)** is only an approximation of the actual value.

Suggest and explain one improvement you would make to the method in **(a)** to increase the accuracy of the experiment.

Use lid to reduce heat loss

OR use 2 Styrofoam cups to reduce heat loss

OR use a pipette or burette for FA 5 to reduce percentage error

OR use magnesium powder so that the reaction can complete faster as there is heat loss while the magnesium ribbon is still reacting

OR use cup with higher walls to reduce acid spray.

[1]

[Total: 13]

3 Investigation of some inorganic reactions.

You are to perform the tests described in **Table 3.1** and record your observations in the table.

In all tests, the reagents should be added gradually until no further change is observed unless you are instructed otherwise. You should indicate clearly at which stage in a test a change occurs, recording your observations alongside the relevant tests.

The volumes given below are approximate and should be estimated rather than measured, unless otherwise stated. Test and identify any gases evolved. If there is no observable change, write **no observable change**.

No additional or confirmatory tests for ions present should be attempted.

FA 7 and **FA 8** are solutions each containing one cation and one anion.

Use a 1 cm depth of **FA 7** or **FA 8** in a test-tube to carry out the following tests on the two solutions and record your observations.

Table 3.1

	Test	*Observations*	
		FA 7	**FA 8**
(a)	Add aqueous sodium hydroxide.	No observable change OR colourless solution OR no precipitate. **[½]**	White precipitate formed which is soluble in excess NaOH(aq). **[½]**
(b)	Add aqueous ammonia.	No observable change OR colourless solution OR no precipitate. **[½]**	White precipitate formed which is insoluble in excess NH_3(aq). **[½]**
(c)	**You do not need to carry out these tests.** Add 1 cm depth of dilute hydrochloric acid, then transfer the mixture into a boiling tube and warm gently.	**Brown gas evolved which turns damp blue litmus paper red.**	**No observable change.**
(d)	Add two or three **drops** of **acidified** aqueous potassium manganate(VII).	Purple $KMnO_4$ decolourises OR brown ppt is formed. [½]	Purple $KMnO_4$ remains. **[½]**

	Test	*Observations*	
		FA 7	**FA 8**
(e)	Add 1 cm depth of aqueous barium nitrate, then add dilute hydrochloric acid.	No observable change OR colourless solution OR no gas with HC*l* OR no precipitate. **[½]**	White precipitate formed which is insoluble in excess HC*l*(aq). **[½]**

Identify as many of the ions present in **FA 7** and **FA 8** as possible from your observations. If you are unable to identify any of the ions from your observations, write 'unknown' in the space.

	FA 7	**FA 8**
Cation	Unknown OR Ba^{2+} **[1]**	Al^{3+} **[1]**
Anion	NO_2^- **[1]**	SO_4^{2-} **[1]**

[8]

(f) **FA 9** is a solid with an anion containing the same element as one of the anions in either **FA 7** or **FA 8** but in a different oxidation state. Relevant anions are listed in the Qualitative Analysis Notes on page 18.

Place a spatula of **FA 9** in a test tube and add 2 cm depth of deionised water. Shake to dissolve the solid and make a solution of **FA 9**.

(i) Select reagents to test whether the anion in **FA 9** contains the same element as the anion in **FA 7**, or in **FA 8**.

Carry out your test(s) on the solution of **FA 9** and record your observations and conclusions below.

	Test	*Observations*	*Conclusion*
(f)	Warm with Zn / Al foil and NaOH(aq) and test the gas with damp red litmus paper. **[1]**	No observable change OR no gas evolved **[½]**	Not nitrate / same element as FA 7 **[½]**
(g)	Add $Ba(NO_3)_2$(aq), followed by adding HC*l*(aq) or HNO_3(aq) OR add two or three **drops** of **acidified** aqueous potassium manganate(VII) **[1]**	White precipitate observed after adding $Ba(NO_3)_2$(aq), which is soluble in excess acid. OR purple $KMnO_4$ decolourises **[1]**	FA 9 contains SO_3^{2-} / sulfite. **[1]**

[5]

[Total : 13]

4 Planning (P)

FB 1 is a Group 2 metal chloride hydrate, MCl_2. n H_2O. In order to determine the identity of M and the value of n, two different methods – volumetric and gravimetric analysis were carried out. The volumetric analysis depends on an application of the principles of solubility product, while gravimetric analysis is used to measure the mass loss on forming the anhydrous metal chloride.

(a) Volumetric analysis

The molar mass of the metal chloride hydrate, MCl_2. n H_2O can be determined by titrating a solution of **FB 1** with aqueous silver nitrate. The equation for this reaction is shown below.

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

Data for use in this question are given in **Table 1**.

Table 1

	colour	**K_{sp} (at 25°C)**
$AgCl$	white	2.02×10^{-10} mol^2 dm^{-6}
Ag_2CrO_4	red	3.01×10^{-12} mol^3 dm^{-12}

FB 1 is the Group 2 metal chloride hydrate, MCl_2. n H_2O(s)

FB 2 is 0.200 mol dm^{-3} silver nitrate

FB 3 is neutral chromate indicator, K_2CrO_4(aq)

The following procedure was carried out.

Step 1 Weigh 6.002 g of **FB 1** using a weighing bottle and transfer the contents completely into a 100 cm^3 beaker.

Step 2 Add deionised water to the beaker and stir until the solid has dissolved.

Step 3 Transfer the contents and washings carefully into a 250 cm^3 volumetric flask and top up the volumetric flask with deionised water to the 250 cm^3 mark. This solution is labelled as **FB 4**.

Step 4 Use a pipette to transfer 25.0 cm^3 of **FB 4** into a conical flask.

Step 5 Titrate the solution with **FB 2** using 10 drops of **FB 3** indicator.

(i) Aqueous silver nitrate is separately added to solutions containing 0.1 mol dm^{-3} $Cl^-(aq)$ and 0.01 mol dm^{-3} $CrO_4^{2-}(aq)$.

Calculate the concentration of Ag^+ that must be present to cause the precipitation of

AgCl: for precipitation to occur, $[Ag^+][Cl^-] > K_{sp}$ (AgCl)

$$\min [Ag^+] = \frac{2.02 \times 10^{-10}}{0.1}$$

$$= \mathbf{2.02 \times 10^{-9}\ mol\ dm^{-3}}\ \mathbf{[1]}$$

Ag_2CrO_4: for precipitation to occur, $[Ag^+]^2 [CrO_4^{2-}] > K_{sp}$ (Ag_2CrO_4)

$$\min [Ag^+] = \sqrt{\frac{3.01 \times 10^{-12}}{0.01}}$$

$$= \mathbf{1.73 \times 10^{-5}\ mol\ dm^{-3}}\ \mathbf{[1]}$$

[2]

(ii) Using the data given in **Table 1** and your answer to **(a)(i)**, suggest what you would see at the beginning of the titration and at the end–point, and explain why $K_2CrO_4(aq)$ can be used as an indicator in this titration.

Beginning of titration: white ppt; End point: red ppt (accept pink) **[1]**

$[Ag^+]$ needed to precipitate Ag_2CrO_4 is much more than that for AgCl and this allows the Cl^- ions present to be sufficiently precipitated out before the Ag_2CrO_4 is precipitated out. **[1]**

(Distinct colour change)

[2]

(iii) **FB 3** contains a neutral solution of $K_2CrO_4(aq)$. Suggest a reason why the titration cannot be carried out in an alkaline condition.

Precipitate of AgOH (brown ppt) will be produced and more $AgNO_3$ will be required and this affects the accurate determination of the end point of the titration / colour change will not be distinct.

[1]

(iv) In the experiment, 24.60 cm^3 of **FB 2** was used to completely react with 25.0 cm^3 of **FB 4**.

Outline how this result can be used to determine the molar mass of MCl_2. n H_2O.

Amount of $AgNO_3$ = $\frac{24.60}{1000} \times 0.200$ = 0.00492 mol

Amount of Cl^- in 25.0 cm^3 = 0.00492 mol

Amount of Cl^- in 250 cm^3 = **0.0492 mol** **[1]**

Amount of $MCl_2.nH_2O$ in 250 cm^3 = $\frac{0.0492}{2}$ = **0.0246 mol** **[1]**

Molar mass of $MCl_2.nH_2O$ = $\frac{6.002}{0.0246}$ = **244 g mol^{-1} [1 with units] (no need 1 dp)**

[3]

(b) Gravimetric analysis

Write a plan to measure the mass loss on forming the anhydrous metal chloride salt, so that the identity of *M* and the value of *n* can be determined.

You may assume that you are provided with the following.

- 5 g of solid **FB 1**
- the apparatus normally found in a school or college laboratory

Your plan should include:

- details of the experimental procedure; **[1]**
- details to ensure reliability of the results; **[1]**
- how your results would be recorded and presented in a table form; **[1]**
- how you would use your results and answer to **(a)(iv)** to determine the value of *n* and the identity of *M*. **[1]**

You may find it useful to use the letters *a*, *b*, *c* etc. to represent each **measurement** of mass.

Procedure

1. Weigh a clean and dry boiling tube (or crucible).
2. Add about 5 g of **FB 1** into the boiling tube and weigh accurately the boiling tube and its contents.
3. Heat the boiling tube gently for about 1 minute and then strongly for a further four minutes.
4. Allow the boiling tube to cool before re–weighing the boiling tube and its contents.
5. Repeat the heating, cooling and weighing process until the mass of the boiling tube and its contents is consistent within 0.01 g in difference.

Results

Mass of boiling tube / g	*a*
Mass of boiling tube + **FA 1** / g	*b*
Mass of **FA 1** used / g	*b* – *a*
Mass of boiling tube + residue (after 1st heating) / g Mass of boiling tube + residue (after 2nd heating) / g	c_1 c_2
Mass of residue / g	c_2 – *a*

Calculation

Mass of H_2O lost = $b - c_2$

= mass of **FA 1** used – mass of residue (alternative)

= $(b - a) - (c_2 - a)$

Amount of H_2O = $\frac{b - c_2}{18.0}$

Amount of $MCl_2.nH_2O$ = $\frac{b - a}{244}$

$\frac{Amount\ of\ H_2O}{Amount\ of\ MCl_2 \cdot nH_2O} = \frac{n}{1}$

$n = \frac{244(b - c_2)}{18(b - a)}$

$M = 244 - 2(35.5) - 18[\frac{244(b - c_2)}{18(b - a)}]$

$= 173 - \frac{244(b - c_2)}{(b - a)}$

[1/2]: logical sequence of steps for experimental procedure

(weigh → heat → cool → reweigh)

[1/2]: use correct glassware

[1/2]: heating, cooling and weighing process

[1/2]: until consistent mass of $\pm$ 0.01 g

[1]: show clear outline to obtain value of *n* and identity of *M*

[1]: correct table

[4]

[Total: 12]

Qualitative Analysis Notes

[ppt. = precipitate]

(a) Reactions of aqueous cations

cation	*reaction with*	
	NaOH(aq)	**NH_3(aq)**
aluminium, Al^{3+}(aq)	white ppt. soluble in excess	white ppt. insoluble in excess
ammonium, NH_4^+ (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), Cr^{3+}(aq)	grey-green ppt. soluble in excess giving dark green solution	grey-green ppt. insoluble in excess
copper(II), Cu^{2+}(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), Mn^{2+}(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, Zn^{2+}(aq)	white ppt. soluble in excess	white ppt. soluble in excess

(b) Reactions of anions

anion	*Reaction*
carbonate, CO_3^{2-}	CO_2 liberated by dilute acids
chloride, Cl^-(aq)	gives white ppt. with Ag^+(aq) (soluble in NH_3(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, NO_3^- (aq)	NH_3 liberated on heating with OH^-(aq) and Al foil
nitrite, NO_2^- (aq)	NH_3 liberated on heating with OH^-(aq) and Al foil; NO liberated by dilute acids (colourless NO → (pale) brown NO_2 in air)
sulfate, SO_4^{2-} (aq)	gives white ppt. with Ba^{2+}(aq) (insoluble in excess dilute strong acids)
sulfite, SO_3^{2-} (aq)	SO_2 liberated with dilute acids; 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	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	turns aqueous acidified potassium manganate(VII) from purple to colourless

(d) Colour of halogens

halogen	*colour of element*	*colour in aq. 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, I_2	black solid / purple gas	brown	Purple