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CLASS **21S** _____**JURONG PIONEER JUNIOR COLLEGE**
JC2 PRELIMINARY EXAMINATION 2022**CHEMISTRY****Higher 2**

Paper 4 Practical

9729/04**16 August 2022****2 hours 30 minutes****Candidates answer on the Question paper.**

Additional Materials: As listed in the Confidential Instructions

READ THESE INSTRUCTIONS FIRST

Write your name, class and exam index number on all the work you hand in.

Give details of the practical shift and laboratory where appropriate, in the boxes provided.

Write in dark blue or black pen.

You may use a HB pencil for any diagrams, 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 18 and 19.

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****2****3****4****Total**This document consists of **19** printed pages.

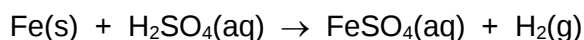
Answer **all** the questions in the spaces provided.

1 Determination of the percentage by mass of iron in a sample of iron wire

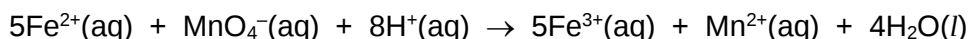
For
Examiner's
Use

Iron wire contains impurities. You will investigate the percentage by mass of iron in a sample of iron wire.

A sample of iron wire is reacted with an excess of sulfuric acid to produce a solution of iron(II) sulfate.



You will titrate the solution of iron(II) sulfate with potassium manganate(VII) of known concentration to determine the amount of iron(II) ions present and hence the percentage by mass of iron in the wire. You may assume the impurities do not form any products that react with potassium manganate(VII).



FA 1 is a solution of FeSO_4 prepared by reacting 3.03 g of iron wire with sulfuric acid to make 500 cm^3 of solution.

FA 2 is 3.16 g dm^{-3} potassium manganate(VII), KMnO_4 .

FA 3 is 1.00 mol dm^{-3} sulfuric acid, H_2SO_4 .

(a) Method

1. Fill the burette labelled "**FA 2**" with **FA 2**.
2. Pipette 25.0 cm^3 of **FA 1** into a conical flask.
3. Use the 50 cm^3 measuring cylinder to transfer 25 cm^3 of **FA 3** to the same conical flask.
4. Titrate this mixture with **FA 2** until a permanent pale pink colour is obtained.
5. Record your titration results, to an appropriate level of precision, in the space provided below.
6. Repeat points 2 to 5 until consistent results are obtained.

Results

[3]

- (b) From your titrations, obtain a suitable volume of **FA 2**, $V_{\text{FA 2}}$, to be used in your calculations. Show clearly how you obtained this volume.

$V_{\text{FA 2}} = \dots\dots\dots$ [3]

- (c) (i) Calculate the amount of KMnO_4 present in the volume of **FA 2**, $V_{\text{FA 2}}$, calculated in (b).

[Ar: O, 16.0; K, 39.1; Mn, 54.9]

amount of MnO_4^- present = [1]

- (ii) Calculate the amount of Fe^{2+} ions in 25.0 cm^3 of **FA 1**.

amount of Fe^{2+} ions = [1]

- (iii) Calculate the percentage by mass of iron in the sample of iron wire.

[Ar: Fe, 55.8]

percentage by mass of iron in the wire = [3]

(d) The maximum error in any single burette reading is $\pm 0.05 \text{ cm}^3$.

- (i) State the maximum possible error in the volume run from the burette recorded in any titration.

..... [1]

- (ii) Explain how an error of $(-0.10) \text{ cm}^3$ may arise when using a burette to obtain an individual titre value.

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.....
..... [1]

- (e) A student suggested that when a piece of iron wire was dissolved in a known volume and concentration of sulfuric acid, the number of moles of iron that reacted with the acid could be determined by working out how much acid was left after the reaction. The amount of excess acid could be determined by titrating the mixture with a known concentration of sodium hydroxide.

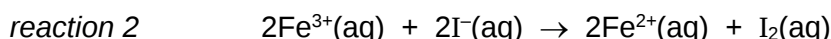
Explain whether the student was correct.

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..... [1]

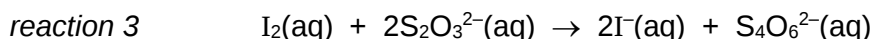
[Total: 14]

2 Investigation of the kinetics of the reaction between $\text{I}^-(\text{aq})$ and $\text{Fe}^{3+}(\text{aq})$

Iron(III) ions, $\text{Fe}^{3+}(\text{aq})$, and iodide ions, $\text{I}^-(\text{aq})$, react to give iodine, $\text{I}_2(\text{aq})$, in the presence of acid.



If acidified iron(III) chloride and potassium iodide are mixed in the presence of starch indicator and sodium thiosulfate, the iodine liberated reacts immediately with the thiosulfate ions and is reduced back to iodide ions.



When all the thiosulfate ions have reacted, the iodine which continues to be produced then turns the starch indicator blue-black. The rate of *reaction 2* may be determined by timing how long it takes for the reaction to turn blue-black.

You are to investigate how the rate of reaction is affected by changing the concentration of the iodide ions.

FA 4 is $0.0500 \text{ mol dm}^{-3}$ potassium iodide, $\text{KI}(\text{aq})$

FA 5 is $0.0500 \text{ mol dm}^{-3}$ acidified iron(III) chloride, $\text{FeCl}_3(\text{aq})$,

FA 6 is $0.00500 \text{ mol dm}^{-3}$ sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3(\text{aq})$

FA 7 is starch indicator.

Read through the whole method carefully and prepare a table for your results on page 6 before starting any practical work.

(a) Method**Experiment 1**

1. Fill the burette labelled "**FA 4**" with **FA 4**.
2. Run 20.00 cm^3 of **FA 4** into the 100 cm^3 beaker.
3. Use the 50 cm^3 measuring cylinder to add the following to the same 100 cm^3 beaker:
 - 20.0 cm^3 of **FA 6**
 - 10.0 cm^3 of **FA 7**
4. Place the beaker on the white tile and stir the contents of the beaker using a glass rod.
5. Use the 25 cm^3 measuring cylinder to measure 10.0 cm^3 of **FA 5**.
6. Add this **FA 5** into the same 100 cm^3 beaker and start timing immediately at the instant of mixing.
7. Stir once. The reaction mixture will turn *light brown* and then *pale yellow* before a sudden intense blue-black colouration is observed. Stop timing as soon as the solution turns blue-black.
8. Record this reaction time, **t**, to the nearest 0.1 s.
9. Wash out the beaker and dry it with a paper towel.

Experiment 2

1. Fill another burette with deionised water.
2. Run 7.00 cm³ of **FA 4** into the 100 cm³ beaker.
3. Run 13.00 cm³ of deionised water into the beaker containing **FA 4**.
4. Use the 50 cm³ measuring cylinder to add the following to the same 100 cm³ beaker:
 - 20.0 cm³ of **FA 6**
 - 10.0 cm³ of **FA 7**
5. Place the beaker on the white tile and stir the contents of the beaker using a glass rod.
6. Use the 25 cm³ measuring cylinder to measure 10.0 cm³ of **FA 5**.
7. Add the **FA 5** to the same 100 cm³ beaker and start timing immediately at the instant of mixing.
8. Stir once and stop timing as soon as the solution turns blue-black.
Note: This may take **more than 2 minutes**.
9. Record this reaction time, **t**, to the nearest 0.1 s.
10. Wash out the beaker and dry it with a paper towel.

Experiments 3 – 5

Carry out three further experiments to investigate how the reaction time, **t**, changes with different volumes of potassium iodide, **FA 4**.

The combined volume of **FA 4** and deionised water must always be 20.00 cm³.

Do not use a volume of **FA 4** that is less than 6.00 cm³.

Record all your results in a single table in the space provided below. You should include the experiment number, volume of **FA 4**, $V_{\text{FA 4}}$, volume of water, V_{water} , and the reaction time, **t**.

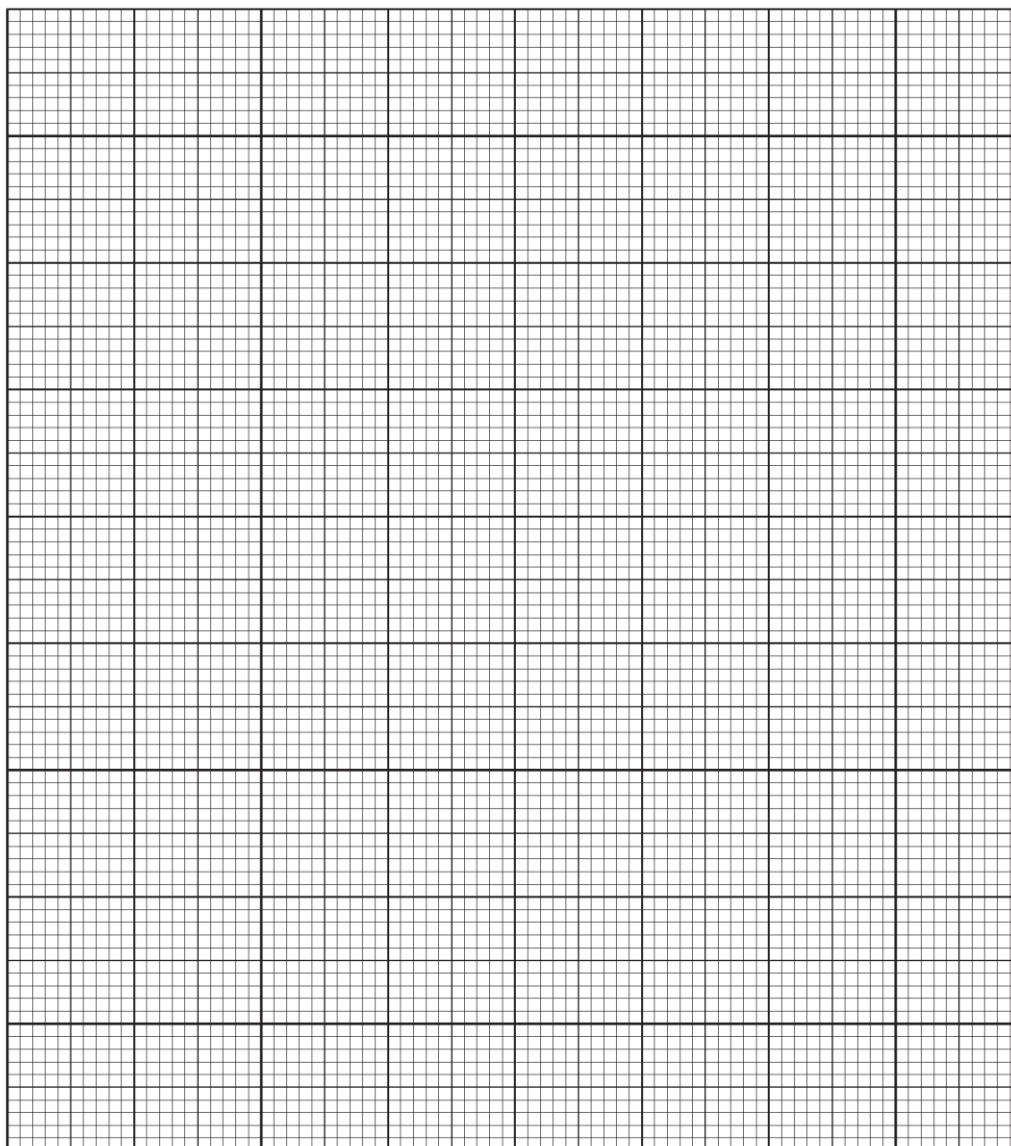
You should also include the 'rate of reaction' which is to be calculated as shown:

$$\text{rate} = \frac{1000}{\text{reaction time}}$$

For each experiment, the value of the 'rate of reaction' is to be recorded to 3 significant figures.

Results

- (b) (i) On the grid below, plot a graph of rate of reaction (y-axis) against volume of **FA 4** (x-axis). Include the origin, (0,0), in your scale. Circle any points you consider anomalous and draw a line of best fit.



[3]

- (b) (ii) Explain, by referring to your graph, how the rate of reaction is affected by an increase in the concentration of aqueous potassium iodide, **FA 4**.

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[2]

(c) Thiosulfate ions can also react with acid to form sulfur, sulfur dioxide and water.

- (i) Write an ionic equation for the reaction between thiosulfate ions and hydrogen ions in aqueous solution. Include state symbols.

..... [1]

- (ii) A student carries out the same investigation as in (a) but the solutions are mixed in different order. The student places **FA 4** and appropriate volume of deionised water in one beaker and all other reactants in a second beaker. The student then transfers the mixture from the second beaker to the first and starts timing.

Tick the box for the statement you consider correct. Explain your answer.

The student's method is better than in (a). ☐

The two methods are equally good. ☐

The student's method is not as good as that in (a). ☐

Reason:

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..... [1]

- (d) Another student investigates the effect of iron(III) concentration on the rate of this reaction.

The student carries out another experiment, **Experiment 6**, and the rate is compared to that of **Experiment 2**.

In **Experiment 2**, the volumes used were:

reagent	Volume / cm ³
FA 4	7.00
FA 5	10.0
FA 6	20.0
FA 7	10.0
deionised water	13.00

- (i) Suggest the volumes the student could use for **Experiment 6**.

reagent	Volume / cm ³
FA 4	
FA 5	
FA 6	
FA 7	
deionised water	

[1]

- (ii) This student records a time of 140 s for **Experiment 2**.

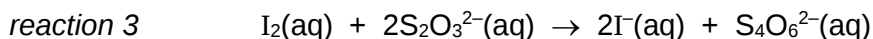
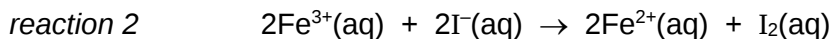
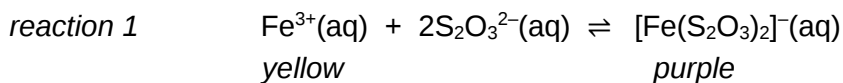
The rate of reaction is directly proportional to the concentration of iron(III) ions.

Suggest how long it would take the reaction mixture proposed for **Experiment 6** in (d)(i) to turn blue-black. Assume that **Experiment 6** is carried out at the same temperature as **Experiment 2**.

Do not carry out **Experiment 6**.

reaction time = s [1]

(e) When **FA 4**, **FA 5** and **FA 6** are mixed, the following three reactions occur.



When **FA 5** was added to the solution prepared in step 6 of **Experiment 1**, the reaction mixture turned *light brown* and then *pale yellow* before the blue–black colour appeared.

The *light brown* colour is due to a mixture of purple $[\text{Fe}(\text{S}_2\text{O}_3)_2]^{-}(\text{aq})$ and yellow $\text{Fe}^{3+}(\text{aq})$ in the reaction mixture. Explain why the *light brown* solution turns *pale yellow* before the appearance of blue-black colour.

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[1]

[Total: 15]

3 Qualitative Analysis

Solid **FA 8** contains the polyatomic anion, X^{2-} , and one cation from the ions listed in the *Qualitative Analysis Notes*.

FA 9 is an aqueous solution containing a sodium salt of the polyatomic anion, X^{2-} .

FA 10 is organic solution containing one functional group.

You will perform tests to

- identify the cation present in **FA 8**
- deduce the nature of **FA 8**
- deduce the identity of X^{2-}
- identify the functional group present in **FA 10**

- (a) (i) Perform the tests described in Table 3.1, and record your observations in the table. Test and identify any gases evolved. If there is no observable change, write **no observable change**.

Table 3.1

tests		observations
1	Place a spatula measure of FA 8 in a test-tube and add 1-2 cm depth of aqueous iron(II) sulfate. Shake the tube well until all FA 8 dissolves.	
2	Place a spatula measure of FA 8 in a boiling tube and add 1 cm depth of dilute nitric acid to dissolve the solid. Then add 1 cm depth of aqueous manganese(II) sulfate and 4 drops of aqueous silver nitrate to act as a catalyst. Heat cautiously to bring the mixture to boiling.	
3	To a spatula measure of FA 8 in a boiling tube, add 1–2 cm depth of aqueous sodium hydroxide. Heat cautiously.	

[3]

(ii) Identify the cation in **FA 8**.

cation:

[1]

(iii) State, with supporting evidences, the nature (*acidic, basic, amphoteric, oxidising or reducing*) of **FA 8**.

nature of **FA 8**:

evidence:

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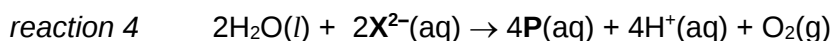
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[2]

(b) (i) **FA 9** is an aqueous solution containing a sodium salt of the polyatomic anion, X^{2-} .

In aqueous solution, X^{2-} reacts very slowly with water to produce anion **P**, H^+ ions and oxygen gas.



Perform the tests described in Table 3.2, and record your observations in the table. Test and identify any gases evolved. If there is no observable change, write **no observable change**.

Table 3.2

tests		observations
1	Test solution FA 9 with Universal Indicator paper.	
2	To 1 cm depth of FA 9 in a boiling tube, add 1 cm depth of aqueous sodium hydroxide and a piece of aluminium foil. Warm the mixture cautiously.	
3	To 1 cm depth of FA 9 in a test-tube, add 1 cm depth of aqueous barium nitrate. Then add dilute nitric acid until no further change occurs.	

[3]

(ii) Identify the anion **P**. Hence, use *reaction 4* equation to deduce the identity of X^{2-} .

Anion **P**: X^{2-} is

[2]

- (c) You are provided with an organic solution **FA 10** which contains one functional group.

Care: FA 10 is flammable. Do not use Bunsen burner for heating. Use the hot water provided if heating is required.

FA 10 gives a positive test with 2,4–dinitrophenylhydrazine.

Devise one other confirmatory test using the bench reagents provided to identify the functional group present in **FA 10**.

Carry out the test. Record details of the test performed and observations made in Table 3.3.

Table 3.3

<i>Confirmatory Test</i>	<i>Observations</i>

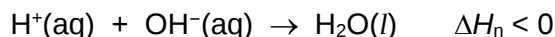
Functional group present in **FA 10**:

[3]

[Total: 14]

4 Planning

Enthalpy change of neutralisation, ΔH_n , is the heat evolved when one mole of water is formed from the reaction of an acid and an alkali.



When aqueous hydrochloric acid, $\text{HCl}(\text{aq})$ is mixed with aqueous sodium hydroxide, $\text{NaOH}(\text{aq})$, the neutralisation reaction releases heat causing a rise in the temperature of the solution.



A series of experiments can be performed where increasing volumes of $\text{HCl}(\text{aq})$ and decreasing volumes of $\text{NaOH}(\text{aq})$ are mixed and the temperature rise, ΔT , for each experiment is determined.

In each of the experiments using different volumes of $\text{HCl}(\text{aq})$ and $\text{NaOH}(\text{aq})$, 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 liberated 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 $\text{HCl}(\text{aq})$ 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 $\text{HCl}(\text{aq})$ and the enthalpy change of neutralisation, ΔH_n , between $\text{HCl}(\text{aq})$ and $\text{NaOH}(\text{aq})$ can be determined.

The volume of $\text{HCl}(\text{aq})$ used should be at least 10.00 cm^3 and the total volume of the reaction mixture should be kept constant at 50.00 cm^3 for all experiments.

- (a) Using the information given, you are required to write a plan to determine the concentration of $\text{HCl}(\text{aq})$ and the enthalpy change of neutralisation, ΔH_n , between $\text{HCl}(\text{aq})$ and $\text{NaOH}(\text{aq})$.

You may assume that you are provided with:

- 250 cm^3 of $\text{HCl}(\text{aq})$ of unknown concentration,
- 250 cm^3 of 1.50 mol dm^{-3} $\text{NaOH}(\text{aq})$,
- 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 $\text{HCl}(\text{aq})$ and $\text{NaOH}(\text{aq})$,
- the procedure you would follow,
- the measurements you would make to allow a suitable ΔT against the volume of $\text{HCl}(\text{aq})$ graph to be drawn.

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[illegible]

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[6]

(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 HCl(aq) needed to just completely neutralise $(50 - V_{\text{neut}}) \text{ cm}^3$ of NaOH(aq) .
- ΔT_{max} , the maximum temperature rise when stoichiometric amount of HCl(aq) and NaOH(aq) reacted.

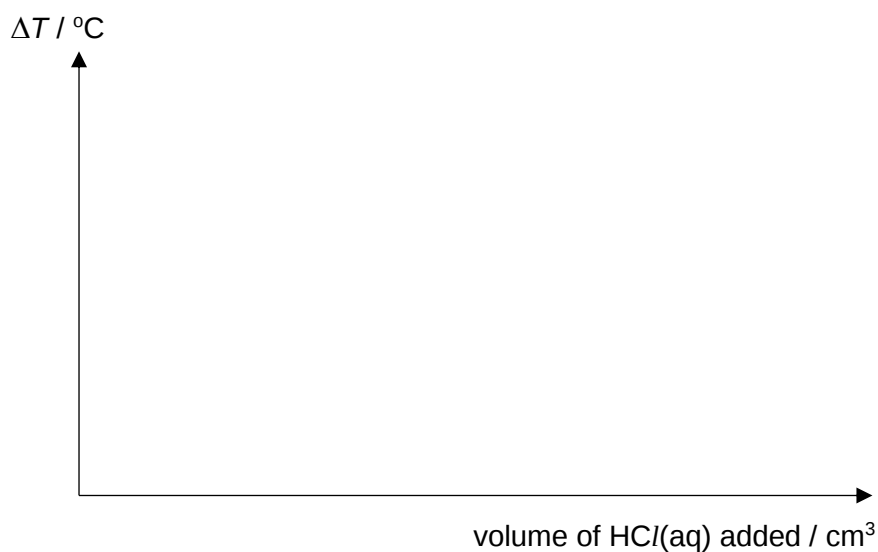


Fig 4.1

[2]

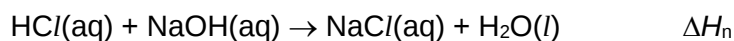
(c) Outline how you would use your answers from (b) to determine

(i) the concentration, in mol dm^{-3} , of HCl(aq) .

Express your results in terms of V_{neut} . It is not necessary to simplify the expression in your final answer.

[2]

(ii) the enthalpy change of neutralisation, ΔH_{n} , for this reaction.



Express your answers in terms of ΔT_{max} and V_{neut} . It is not necessary to simplify the expression in your final answer.

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} .

[2]

[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 $\text{NO} \rightarrow$ (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 on warming with dilute acids; gives white ppt. with Ba^{2+} (aq) (soluble in excess dilute strong acids)

(c) Tests for Gases

gas	Test and test results
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 acidified aqueous 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, I_2	black solid/purple gas	brown	purple