SALTS

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Organizer

Chapter objectives

By the end of this chapter, the learner should be able to:

(a) Define a salt.
(b) State and describe the types of salts.
(c) Identify soluble and insoluble salts.
(d) State the methods of preparing soluble and insoluble salts.
(e) Define the terms: saturated solution, crystallisation, neutralisation and precipitation.
(f) Write ionic equations for the preparation of insoluble salts.
(g) State the effect of heat on salts.
(h) State the uses of some salts.

(20 Lessons)

SALTS

A **salt** is *a substance that is formed when* ***the hydrogen ions in an acid are wholly or partially*** *replaced by* ***a positive ion*****.**

Types of Salts

There are four types of salts namely; normal salts, acid salts, basic salts and double salts.

Normal salts

A **normal salt** *is a salt that* ***does not*** *contain any* ***replaceable hydrogen atom*****.**

Some examples of normal salts are sodium chloride (NaCl), potassium sulphate (K_2SO_4) sodium carbonate (Na_2CO_3) and calcium nitrate ($Ca(NO_3)_2$). These salts are **neutral** in aqueous state.

Acid salts

Acid salts *are salts that* ***contain a replaceable hydrogen atom*****.**

Some examples are sodium hydrogen carbonate ($NaHCO_3$), potassium hydrogen sulphate ($KHSO_4$) and sodium dihydrogen phosphate (NaH_2PO_4). They have **acidic** properties due to the presence of **replaceable hydrogen**.

Basic salts

Basic salts *are salts that contain* ***hydroxyl (OH^-) i****ons.*

It is the presence of hydroxyl ions in these salts that is responsible for the basic properties. Examples are basic magnesium chloride (Mg(OH)Cl, basic lead (II) carbonate ($Pb(OH)_2.PbCO_3$), basic zinc chloride (Zn(OH)Cl) and basic copper (II) carbonate ($CuCO_3$).$Cu(OH)_2$).

Double salts

Double salts *are salts in which there are* ***two different anions or cations.***

Examples are hydrated potassium aluminium sulphate ($KAl(SO_4)_2.12H_2O$) hydrated ammonium iron (II) sulphate ($Fe(NH_4)(SO_4).6(H_2O)$) and trona (Na_2CO_3. $NaHCO_3.2H_2O$).

Solubility of Salts and Bases in Water

Solubility of Salts in Water

Discussion Questions

1. **Which cations form salts that are soluble in water?**
- All potassium, sodium and ammonium salts are soluble in water.

2. **Comment on the solubility of sulphate, chlorides, nitrates and carbonates**
- All nitrates are soluble.
- All sulphates are soluble except those of barium and lead. Calcium sulphate is slightly soluble.

- All chlorides are soluble except lead (II) chloride and silver chloride. Lead (II) chloride is insoluble in cold water but soluble when hot.
- All carbonates are **insoluble** except those of Group I metals and ammonium. Carbonates of aluminium and iron do not exist.

Summary of Solubility of Common Salts

SALTS	SOLUBILITY IN WATER
Carbonates	**All are insoluble except potassium sodium and ammonium carbonate.**
Chlorides	**All are soluble except silver and lead (II) chloride. Lead (II) chloride is soluble in hot water.**
Nitrates	**All are soluble**
Sulphates	**All are soluble except barium sulphate and lead(II) sulphate. Calcium sulphate is slightly soluble.**
	All potassium, sodium and ammonium salts are soluble in water

Solubility of Bases in Water

Discussion Questions

1. **Which oxides are Soluble? What is the nature of the bases formed?**
   - The oxides potassium and sodium are **soluble** in water.
   - The oxides of calcium and magnesium are **slightly soluble**.
   - The solution formed is alkaline and turns red litmus paper blue.

$$\text{Calcium oxide + water} \longrightarrow \text{Calcium hydroxide}$$

$$CaO(s) + 2H_2O(l) \longrightarrow Ca(OH)_2\ (aq)$$

$$\text{Magnesium oxide + water} \longrightarrow \text{Magnesium hydroxide}$$

$$MgO(s) + H_2O(l) \longrightarrow Mg(OH)_2\ (aq).$$

2. **Which Oxides are Insoluble?**
   - The oxides of zinc , aluminium, lead, iron and copper are insoluble in water.

Ammonium oxide does not exist.

3. **Which hydroxides are Soluble? What is the nature of the solutions formed?**
   - Potassium and sodium hydroxides are soluble in water.
   - The hydroxides of calcium and magnesium are slightly soluble.
   - The resulting solutions are **alkaline** and **turn red litmus blue**.

$$\text{Potassium hydroxide + water} \longrightarrow \text{Potassium hydroxide solution}$$

$$KOH(s) + \text{water} \longrightarrow KOH\ (aq)$$

$$\text{Sodium hydroxide + water} \longrightarrow \text{Sodium hydroxide solution}$$

$$NaOH(s) + \text{water} \longrightarrow NaOH(aq)$$

4. Which hydroxides are Insoluble?

- The **hydroxides** of **zinc, aluminium, iron, lead and copper are insoluble in water**.
- Ammonia is a base and dissolves in water to form aqueous ammonia.

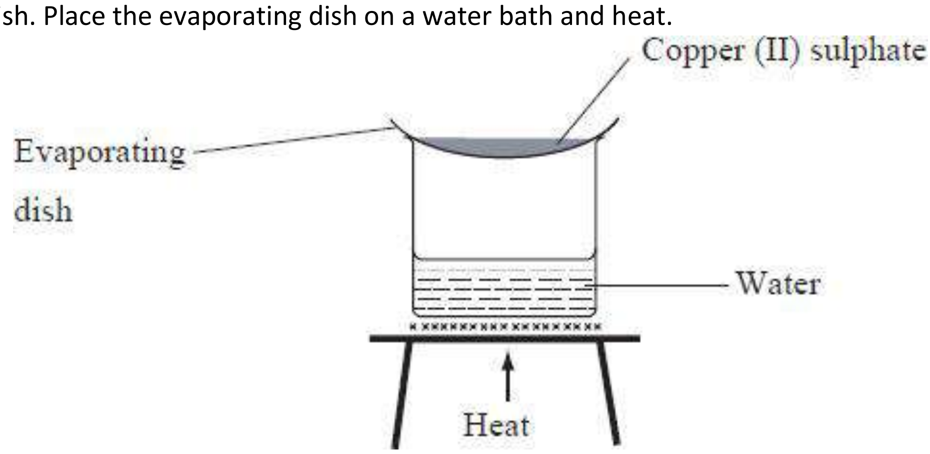
Cation	*Solubility*		*Effect of resulting solution on litmus paper*	
	Oxide	**Hydroxide**	**Oxide**	**Hydroxide**
Potassium	Soluble	Soluble	Turns blue	Turns blue
Sodium	Soluble	Soluble	Turns blue	Turns blue
Calcium	Slightly soluble	Slightly soluble	Turns blue	Turns blue
Magnesium	Slightly soluble	Slightly soluble	Turns blue	Turns blue
Zinc	Not soluble	Not soluble	No effect	No effect
Aluminium	Not soluble	Not soluble	No effect	No effect
Copper	Not soluble	Not soluble	No effect	No effect

How to obtain crystals of a salt from a solution

It is possible to obtain crystals of a salt from a solution by evaporating the solvent.

For example, to obtain crystals of copper (II) sulphate obtained from a solution of copper (II) sulphate, the following procedure is used:

Measure about 20 cm^3 of water in a beaker. Add a spatulaful of copper (II) sulphate crystals in the beaker and stir. Continue adding until it does not dissolve anymore. Decant a portion of the solution into an evaporating dish. Place the evaporating dish on a water bath and heat.



Evaporate the solution until it is about to form crystals. To find out if this point has been reached, dip a clean glass rod into the solution and hold it up in the air to cool. Examine the rod to find out whether the crystals have formed on it. Continue heating until crystals are seen on the glass rod when the rod is dipped in the solution. Allow the solution to cool slowly to form crystals. Observe the crystals with a hand lens.

To obtain larger crystals, cover the evaporating dish with a perforated paper. Leave the set-up undisturbed and observe after 12 hours. Filter off the crystals and dry them between filter papers.

Discussion Questions

1. **How can you tell whether a solution has dissolved as much copper (II) sulphate as possible?**
   When the blue solution becomes **saturated** such that no more of the salt can dissolve and some crystals remain at the bottom of the beaker.

A **saturated solution** *is one that cannot dissolve any more solute at a given temperature*.

1. **What happens as the saturated solution cools?**

   As the hot saturated solution cools, it forms **solid particles with a regular shape.** These solid particles are called **crystals** and the process is referred to as **crystallisation**.

2. **Explain why the solution is not evaporated to dryness.**

   Evaporation is not done to dryness so that larger crystals may form. Slow cooling allows the salt to form large crystals. Otherwise the crystals formed would be small.

3. **Why are the crystals not dried by heating?**

   Crystals of copper (II) sulphate are not dried by heating because heating would drive out water of crystallisation to leave behind a powder, anhydrous copper (II) sulphate.

4. **What is water of crystallization? How important is it?**

   The fixed amount of water incorporated within the crystal structure of a salt is called **water of crystallisaton**.

   Water of crystallisation is important in the formation of crystals for hydrated salts.

   Hydrated salts are salts which contain water of crystallization.

   Examples of hydrated salts include; sodium carbonate decahydrate ($Na_2CO_3.10H_2O$), Iron (II) sulphate heptahydrate ($FeSO_4.7H_2O$).

3. Methods of Preparing Salts

The method chosen for preparation a specific salt may depend on the solubility of the salt in water.

Preparation of Soluble salts

Soluble salts can be prepared in many ways.

Reaction of an acid and a metal.

The reaction of an acid and a metal is suitable for the preparation of soluble salts such as zinc sulphate, magnesium nitrate, zinc chloride and calcium chloride. The metals are reacted with suitable acids.

Equations

$$\text{Zinc + Sulphuric (VI) acid} \longrightarrow \text{Zinc sulphate + Hydrogen gas}$$

$$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$$

$$\text{Magnesium + Nitric acid} \longrightarrow \text{Magnesium nitrate + Hydrogen gas}$$

$$Mg(s) + 2HNO_3(aq) \longrightarrow Mg(NO_3)_2\,(aq) + H_2\,(g)$$

$$\text{Zinc + Hydrochloric acid} \longrightarrow \text{Zinc chloride + Hydrogen gas}$$

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

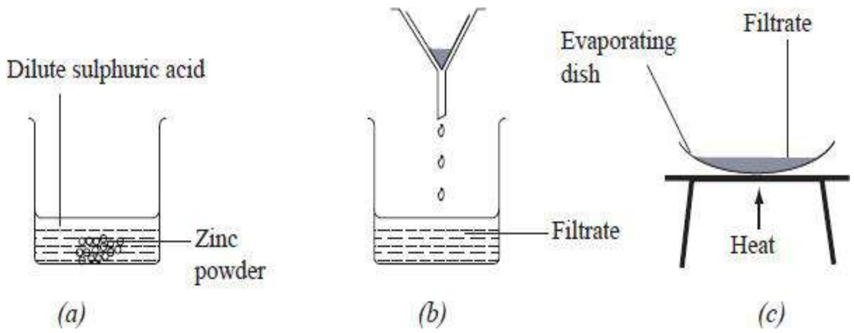
$$\text{Calcium + Hydrochloric acid} \longrightarrow \text{Calcium chloride + Hydrogen gas}$$

$$Ca(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2(g)$$

The procedure below can be used to prepare **zinc (II) sulphate.**

Measure about 20 cm^3 of dilute sulphuric acid (VI) acid and transfer it into a beaker. Add zinc powder a little at a time as you stir with a glass rod. Continue adding zinc powder until it is in excess. Filter the solution and pour the filtrate into an evaporating basin.

Evaporate the filtrate to saturation. To find out if this point has been reached, dip a glass rod into the solution and hold it up in air to cool. If cystals form on the tip of the rod, the solution is ready to form crystals and heating can be stopped to allow the now saturated solution to cool and evaporate slowly.



Questions

1. **Explain why zinc powder is added in excess.**

   Excess zinc is used to ensure all the acid reacts completely.

2. **Why is filtration necessary in this experiment?**

   The unreacted zinc is removed by filtration.

3. **Write an equation for the reaction between zinc powder and dilute sulphuric acid.**

   Zinc reacts with dilute sulphuric (VI) acid to produce zinc sulphate and hydrogen gas. Effervescence stops when all the sulphuric acid has completely reacted with zinc powder.

   Zinc + Sulphuric (VI) acid ⟶ Zinc sulphate + Hydrogen gas

   $$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$$

   The saturated salt solution is allowed to cool for zinc sulphate crystals to form. The process of obtaining salt crystals from a saturated solution by cooling is called **crystallisation**.

Reaction of acids on insoluble bases

The reaction of acids on insoluble bases is another method of preparation of salts, ideal for preparation of soluble salts such as copper (II) sulphate. Other salts that can be prepared in a similar way are lead (II) nitrate, magnesium sulphate and calcium chloride.

Copper (II) oxide + Sulphuric (VI) acid ⟶ Copper sulphate + water

$$CuO(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + H_2O(l)$$

Lead (II) oxide + dilute nitric (V) acid ⟶ Lead nitrate + water

$$PbO(s) + 2HNO_3(aq) \longrightarrow Pb(NO_3)_2(aq) + H_2O(l)$$

Magnesium oxide + dil. sulphuric (VI) acid ⟶ Magnesium sulphate + water

$$MgO(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2O(l)$$

Calcium oxide + dil. hydrochloric acid ⟶ Calcium chloride + water

$$Ca(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l)$$

The **procedure below can be used to prepare copper (II) sulphate.**

Measure about 20 cm^3 of dilute sulphuric (VI) acid and pour it into a glass beaker. Warm the acid in the beaker. Using a spatula, add copper (II) oxide to the warm acid a little at a time while stirring with a glass rod until no more oxide can dissolve. Filter and collect the filtrate. Transfer the filtrate to the evaporating basin. Evaporate the filtrate over a water bath to saturation. Stop heating and allow the saturated solution to cool to form crystals. Dry the crystals between filter papers.

Questions

1. **Why was copper (II) oxide added in excess?**

   Excess copper (II) oxide is used to ensure that all the acid has reacted.

2. **Explain why the acid was warmed before adding copper (II) oxide?**

   Since the reaction between the acid and the oxide is slow, warming speeds up the reaction.

3. **Besides the salt, what is the other product?**

   The reaction between copper (II) oxide and dilute sulphuric (VI) acid produces a salt and water only. This type of reaction is called **neutralisation**.

4. **Why was evaporation done over a water bath?**

   Evaporation of the filtrate is carried out over a water bath to ensure slow evaporation and formation of large crystals.

5. **Write an equation for the reaction between copper (II) oxide and dilute sulphuric acid.**

   Copper (II) oxide + Sulphuric (VI) acid $\longrightarrow$ Copper sulphate + water

   $CuO(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + H_2O(l)$

Neutralization: Reaction of acids on soluble bases

A Neutralisation reaction is a *reaction in which only a salt and water is formed*.

The action of acids on soluble bases is an example of a neutralization reaction and is used to prepare some soluble salts. Sodium chloride, potassium nitrate, ammonium chloride, ammonium sulphate and sodium sulphate are prepared by this method.

The equations given are for the reactions between various acids and bases to produce salts.

Sodium hydroxide + hydrochloric acid $\longrightarrow$ Sodium chloride + water

$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$

Potassium hydroxide + Nitric acid $\longrightarrow$ Potassium nitrate + water

$KOH(aq) + HNO_3(aq) \longrightarrow KNO_3(aq) + H_2O(l)$

Ammonium hydroxide + Hydrochloric acid $\longrightarrow$ Ammonium chloride + water

$NH_4OH\ (aq) + HCl(aq) \longrightarrow NH_4Cl(aq) + H_2O(l)$

Sodium hydroxide + Sulphuric (VI) acid $\longrightarrow$ Sodium sulphate + water

$2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$

The procedure below can be used to prepare sodium chloride.

Measure 25 cm^3 of dilute hydrochloric acid and pour it in a beaker. Dip universal indicator paper into the beaker. Record the pH. Measure 25 cm^3 of sodium hydroxide and pour it in another beaker. Record its pH also. Add 2–3 drops of phenolphthalein into the sodium hydroxide. Pour the hydrochloric acid slowly into the beaker containing sodium hydroxide until the pink colour just disappears. Test the resultant solution with universal indicator paper. Pour about 25 cm^3 of the resultant solution into a clean evaporating basin, evaporate the solution until it is saturated. Allow the saturated solution to cool for crystals to form.

Questions

1. **Explain the role of phenolphthalein in this experiment.**

   Phenolphthalein indicator is pink in alkaline solutions and colourless in acidic solutions. It is used in this reaction to determine when the reaction is over, the point at which just enough of the acid has been added to exactly neutralise the alkali. This point is known as the **end point**.

Universal indicator paper **shows a range of colours** in alkaline and acidic solutions. It is blue in sodium hydroxide and red in hydrochloric acid. It is not very suitable for determining the end point of an acid base reaction.

2. **Write an equation for the reaction that takes place between sodium hydroxide and hydrochloric acid.**

Sodium hydroxide + hydrochloric acid ⟶ Sodium chloride + water

$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

Reaction of acids with carbonates

The reaction of acids with carbonates is another suitable method for preparing soluble salts such as lead nitrate, which can be prepared by reacting lead carbonate with dilute nitric (V) acid.

Lead(II) carbonate + Nitric (IV) acid ⟶ Lead nitrate + Carbon(IV) oxide + Water

$$PbCO_3(a) + 2HNO_3(aq) \longrightarrow Pb(NO_3)_2 + CO_2(g) + H_2O\ (l)$$

Other salts such as calcium chloride, sodium nitrate, zinc nitrate and ammonium sulphate can also be prepared by reacting suitable acids with their corresponding carbonates.

The **procedure below can be used to prepare lead nitrate.**

Measure about 25 cm^3 of dilute nitric (V) acid and pour it into a glass beaker. Warm the acid in the beaker gently. Add lead (II) carbonate to the warm acid, a little at a time and stir. Continue adding the carbonate while stirring until effervescence stops. Stop warming and filter off the unreacted carbonate. Evaporate the filtrate to saturation and allow it to form crystals.

Questions

1. **What are the products of the reaction? Write an equation for the reaction between lead (II) carbonate and nitric acid.**

When lead (II) carbonate is added to the warm nitric acid, there is immediate effervescence. The effervescence is due to production of carbon (IV) oxide gas.

Lead(II) carbonate + Nitric (IV) acid ⟶ Lead nitrate + Carbon(IV) oxide + Water

$$PbCO_3(a) + 2HNO_3(aq) \longrightarrow Pb(NO_3)_2 + CO_2(g) + H_2O\ (l)$$

2. **Write equations to show how the following salts may be prepared from their corresponding carbonates.**

(i) Zinc nitrate.

Zinc carbonate + Nitric (IV) acid ⟶ Zinc nitrate + carbon (IV) oxide + Water

$$ZnCO_3(s) + 2HNO_3(aq) \longrightarrow Zn(NO_3)_2\ (aq) + CO_2(g) + H_2O(l)$$

(ii) Calcium chloride.

Calcium carbonate+ Hydrochloric acid ⟶ Calcium chloride + carbon (IV) oxide + water

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

(iii) Ammonium sulphate.

Ammonium carbonate + Sulphuric (VI)acid ⟶ Ammonium sulphate + Carbon (IV) oxide + Water

$$(NH_4)_2CO_3(s) + H_2SO_4(aq) \longrightarrow (NH_4)_2SO_4(aq) + CO_2(g) + H_2O(l)$$

(iv) Sodium nitrate.

Sodium carbonate + Nitric (IV) acid ⟶ Sodium nitrate + Carbon (IV) oxide + Water

$$Na_2CO_3(s) + 2HNO_3(aq) \longrightarrow 2NaNO_3(aq) + CO_2(g) + H_2O(l)$$

Preparation of salts by Direct Combination of Elements

Salts can also be prepared by reacting a metal with a non-metal. This method is called **direct synthesis** and is suitable for preparing **both soluble and insoluble salts**.

Iron (III) sulphide can be prepared this way by reacting iron with sulphur.

Iron + Sulphur ⟶ Iron (II) sulphide

$$Fe\ (s) + S(s) \longrightarrow FeS(s)$$

Other salts that can be prepared by this method are sodium chloride and iron (III) chloride.

Sodium + Chlorine ⟶ Sodium chloride

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

Iron + Chlorine ⟶ Iron (III) chloride

$$2Fe(s) + 3\ CL_2(g) \longrightarrow 2FeCl_3(s)$$

The procedure below can be used to prepare iron (II) sulphide.

Put a spatulaful of iron filling in a crucible. To the same crucible add a spatulaful of sulphur. Mix them well. Heat the mixture strongly. When the reaction is complete, allow the products to cool.

Questions

1. **What is the colour of:**
   - **(i) Iron filings?**
     The colour of iron is grey
   - **(ii) Sulphur?**
     The colour of sulphur is yellow
   - **(iii) The product?**
     The colour of the product is black. The black solid formed is iron (II) sulphide.
2. **What was observed when the mixture was being heated?**
   When a mixture of sulphur and iron fillings is strongly heated, it glows red even when the source of heat is removed. This shows that the reaction between iron and sulphur produces heat.
3. **Write an equation between iron filings and sulphur.**

Iron + Sulphur ⟶ Iron (II) sulphide

$$Fe\ (s) + S(s) \longrightarrow FeS(s)$$

Preparation of insoluble salts

Precipitation reactions are suitable for preparing insoluble salts.

Reactions in which solids are formed from aqueous solutions are called **precipitation reactions.** The solids formed are referred to as **precipitates.**

In a precipitation reaction, the two reactants must be **soluble salts** and **one of the products must be insoluble.**

In these reactions, the metal ions of the salts simply exchange their anions. This type of reaction is called **double decomposition.**

In precipitation reactions, some ions undergo changes in their physical state ie from solid to aqueous or aqueous to solid. However, the state of other ions remain unchanged.

The ions that remain unchanged during a chemical reaction are called **spectator ions** and are omitted when writing **ionic equations.**

The equation obtained by writing only those ions that undergo change during a chemical reaction is called an **ionic equation.**

Examples of precipitation reactions are:

Barium nitrate + Zinc sulphate ⟶ Barium sulphate + Zinc nitrate

$Ba(NO_3)_2(Aq) + ZnSO_4(aq) \longrightarrow BaSO_4(s) + Zn(NO_3)_2(aq)$

$Ba^{2+}(aq) + 2NO^-_3(aq) + Zn^{2+}(aq) + SO^{2-}_4 \longrightarrow BaSO_4(s) + Zn^{2+}(aq) + 2NO^-_3(g)$

Ionic equation: $Ba^{2+}(aq) + SO^{2-}_4\ (aq) \longrightarrow BaSO_4(s)$

Lead nitrate + Potassium iodide ⟶ Lead iodide + Potassium nitrate.

$Pb(NO_3)_2(aq) + 2KI(aq) \longrightarrow PbI_2(s) + 2KNO_3(aq)$

$Pb^{2+}(aq) + 2NO^-_3(aq) + 2K^+(aq) + 2I^-(aq) \longrightarrow PbI_2(s) + 2K^+(aq) + 2NO^-_3(aq)$

Ionic equation: $Pb^{2+}(aq) + 2I^-(aq) \longrightarrow PbI_2(s)$

Silver nitrate + Sodium chloride ⟶ Silver chloride + Sodium nitrate

$AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$

$Ag^+(aq) + NO^-_3(aq) + Na^+ + Cl^-(aq) \longrightarrow AgCl(s) + Na^+\ (aq) + NO^-_3(aq)$

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

Copper chloride + Sodium carbonate ⟶ Copper carbonate + Sodium chloride

$CuCl_2(aq) + Na_2CO_3(aq) \longrightarrow CuCO_3(s) + 2NaCl(aq)$

$Cu^{2+}(aq) + 2Cl^-(aq) + 2Na^+(aq) + CO_3^{2-}(aq) \longrightarrow CuCO_3(s) + 2Cl^-(aq) + 2Na^+(aq)$

Ionic equation: $Cu^{2+}(aq) + CO^{2-}_3(aq) \longrightarrow CuCO(s)$

The procedure below can be used to prepare lead (II) sulphate

Put 10 cm^3 of lead (II) nitrate in a beaker. To the same beaker, add excess magnesium sulphate solution. Stir the solution using a glass rod. Let the solid settle then decant the liquid. Wash the solid with distilled water. Filter and dry the solid between filter papers.

Questions

1. **Which ions are present in the reactants?**

The ions present in the reactants are lead (Pb^{2+}), and nitrate (NO_3^-) ions from lead (II) nitrate; magnesium (Mg^{2+}) and sulphate (SO_4^{2-}) ions from magnesium sulphate.

2. **What observations are made when lead (II) nitrate and magnesium sulphate solutions are mixed?**

When lead (II) nitrate and magnesium sulphate solutions are mixed, a white solid is formed. When the two salt solutions react, lead sulphate and magnesium nitrate salts are formed.

In this reaction, the metal ions of the salts simply exchange their anions. Lead sulphate is formed as a white solid (precipitate).

3. **Write an equation for the reaction between lead (II) nitrate and magnesium sulphate.**

Lead (II) nitrate + Magnesium sulphate $\longrightarrow$ Lead (II) sulphate + Magnesium nitrate

$$Pb(NO_3)_2(aq) + MgSO_4(aq) \longrightarrow PbSO_4(s) + Mg(NO_3)_2(aq)$$

$$Pb^{2+}(aq) + 2NO^-_3(aq) + Mg^{2+}(aq) + SO^{2-}_4(aq) \longrightarrow PbSO_4(s) + Mg^{2+}(aq) + 2NO^-_3(aq)$$

4. **(i) Name the solid formed in 3 above.**

Lead sulphate

(ii) Which ions react to form the solid.

The ions in the salt solutions that react to form lead sulphate are lead ions (Pb^{2+}) and sulphate ions (SO_4^{2-}). These ions are in aqueous state in the begining of the reaction and end up in the solid state at the end of the reaction. The ions undergo a change in their physical state.

(iii) Write an equation using the ions that form the solid.

$$Pb^{2+}(aq) + SO^{2-}_4(aq) \longrightarrow PbSO_4(s)$$

Magnesium ions (Mg^{2+}) and nitrate ions (NO^-_3) are the spectator ions since they are in aqueous state in the begining and at the end of the reaction, they remain unchanged through out the reaction.

Effects of exposing salts to the atmosphere

Salts undergo some changes after some time when left in the open.

1. Salts that **absorb water from the atmosphere and form solutions** are said to be **deliquescent** and the process is called **deliquescence.**
   When **anhydrous calcium chloride** is exposed to the atmosphere overnight, it forms a colourless solution. The **shiny black crystals of iron (III) chloride** also form a **yellow solution.** This is because they absorb so much water from the atmosphere that they form solutions. Other deliquescent substances are **sodium hydroxide, calcium chloride, iron (II) chloride, potassium hydroxide and zinc chloride.**

2. Salts that **absorb water from the atmosphere but do not form solutions** are said to be **hygroscopic** and the process is called **hygroscopy**.
   When **anhydrous copper (II) sulphate** is exposed to the atmosphere, the **white solid turns blue and becomes damp.** This is because it absorbs water from the atmosphere but does not dissolve. Examples of **hygroscopic salts** are **anhydrous cobalt (II) chloride, potassium nitrate** and **common table salt.**
   NB: Pure sodium chloride **is not hygroscopic**, but common table salt is hygroscopic **because it contains other salts** such as **magnesium chloride** which makes it behave like a hygroscopic salt. It is the magnesium chloride, that absorbs water and makes the salt damp.

3. Salts which **lose some of their water of crystallisation** when exposed to the atmosphere are said to be **efflorescent**. The process is called **efflorescence**.
   When crystals of **hydrated sodium carbonate ($Na_2CO_3.10H_2O$)** are exposed to the atmosphere over-night a white powder is formed. This, is because the crystals lose some water of crystallisation to form a white powder, **sodium carbonate monohydrate. ($Na_2CO_3.H_2O$)**. Other examples of efflorescent salts include **iron(II) sulphate heptahydrate, ($FeSO_4.7H_2O$) and Sodium sulphate deca-hydrate ($Na_2SO_4.10H_2O$)**

Uses of Salts

Salts are widely used. Some of the uses are:

1. Fertilizers contain sodium nitrate, potassium sulphate and ammonium salts such as ammonium nitrate, ammonium sulphate and ammonium phosphate.
2. Sodium chloride is used as a food additive while sodium hydrogen carbonate is used in baking powder.
3. Calcium sulphate (Plaster of Paris) is used in hospitals on people with fractures or dislocations.
4. Calcium chloride is used in road surfacing to keep the road moist due to its deliquescent nature, moisture absorbs dust.
5. Calcium chloride is used in the extraction of sodium metal by electrolysis.
6. Potassium nitrate is used for making fireworks and gunpowder.
7. Sodium carbonate is used in the softening of hard water, making of glass and detergents.
8. Many salts are used for de-frosting snow in cold countries during winter by lowering its freezing point. Calcium chloride is applied on roads to melt snow.

6. Action of heat on salts

Heat energy can bring about changes in substances when applied to them. Some salts undergo thermal decomposition when they are heated.

All ammonium salts decompose on heating.

Action of Heat on Carbonates

Discussion Questions

1. **Which carbonates are not affected by heat? Explain.**

   Pure **carbonates of sodium and potassium** are not affected by heat. Carbonates of these metals are **stable because of the high position of the metals in the reactivity series**.

   If hydrated, the salts only lose their water of crystallisation. The water vapour condenses on the sides of the test tube as a colourless liquid. The residue is a white solid.

$$\text{Sodium carbonate decahydrate} \xrightarrow{heat} \text{Sodium carbonate} + \text{Water of crystallisation}$$

$$Na_2CO_3.10H_2O(s) \xrightarrow{heat} Na_2CO_3(s) + 10H_2O(l)$$

2. **Write equations for the reactions which occur when some carbonates are heated.**

The other metal carbonates decompose on heating to give off a colourless gas and the corresponding metal oxide. On testing the gas with calcium hydroxide, a white precipitate is formed confirming the gas to be carbon (IV) oxide.

The general equation is;

$$\text{Metal Carbonate} \xrightarrow{heat} \text{Metal oxide} + \text{Carbon (IV) oxide}$$

$$\text{Calcium carbonate} \xrightarrow{heat} \text{Calcium oxide} + \text{Carbon (IV) oxide}$$

$$CaCO_3(s) \xrightarrow{heat} CaO(s) + CO_2(g)$$

$$\text{Zinc Carbonate} \xrightarrow{heat} \text{Zinc oxide} + \text{Carbon (IV) oxide}$$

$$ZnCO_3(s) \xrightarrow{heat} ZnO(s) + CO_2(g)$$

$$\text{Lead (II) carbonate} \xrightarrow{heat} \text{Lead (II) oxide} + \text{Carbon (IV) oxide}$$

$$PbCO_3(s) \xrightarrow{heat} PbO(s) + CO_2(g)$$

$$\text{Copper (II) carbonate} \xrightarrow{heat} \text{Copper (II) oxide} + \text{Carbon (IV) oxide}$$

$$CuCO_3(s) \xrightarrow{heat} CuO(s) + CO_2(g)$$

3. **What is the relationship between ease of decomposition of a metallic carbonate and the position of the metal in the reactivity series?**

   The **position of a metal i**n the reactivity series **determines the ease of decomposition of its carbonate.**

Ammonium carbonate decomposes to form ammonia gas, carbon (IV) oxide and steam. Presence of ammonia is confirmed by the turning of moist red litmus paper to blue.

$$\text{Ammonium carbonate} \xrightarrow{heat} \text{Ammonia} + \text{Carbon (VI) oxide} + \text{Steam}$$

$$(NH_4)_2CO_3(s) \xrightarrow{heat} 2NH_3(g) + CO_2(g) + H_2O(g)$$

Hydrogen carbonates of metals high in the reactivity series decompose on heating to produce the corresponding metal carbonates, carbon (IV) oxide and water.

$$\text{Sodium hydrogen carbonate} \xrightarrow{heat} \text{Sodium carbonate} + \text{water} + \text{Carbon (IV) oxide}$$

$$2NaHCO_3(s) \xrightarrow{heat} Na_2CO_3(s) + H_2O(l) + CO_2(g)$$

$$\text{Potassium hydrogen carbonate} \xrightarrow{heat} \text{Potassium carbonate} + \text{Water} + \text{Carbon (IV) oxide}$$

$$2KHCO_3(s) \xrightarrow{heat} K_2CO_3(s) + H_2O(g) + CO_2(g)$$

The hydrogen carbonates of calcium and magnesium exist only in solution whereas hydrogen carbonates of aluminium and iron **do not exist.**

Action of Heat on Nitrates

All nitrates decompose on heating.

Discussion Questions

1. **Identify the products when each nitrate is heated.**
2. **Write equations for the reactions that occur when the metal nitrates were heated.**
3. **What is the general trend of the action of heat on nitrates?**

Potassium nitrate and sodium nitrate decompose to form oxygen and a white residue which is potassium nitrite and sodium nitrite respectively.

$$\text{Potassium nitrate} \xrightarrow{heat} \text{Potassium nitrite} + \text{Oxygen}$$

$$2KNO_3(s) \xrightarrow{heat} 2KNO_2 + O_2(g)$$

$$\text{Sodium nitrate} \xrightarrow{heat} \text{Sodium nitrite} + \text{Oxygen}$$

$$2NaNO_3(s) \xrightarrow{heat} 2NaNO_2(s) + O_2(g)$$

The nitrates of calcium, zinc, lead and copper decompose on heating to form the metal oxide, nitrogen (IV) oxide and oxygen gas.

$$\text{Calcium nitrate} \xrightarrow{heat} \text{Calcium oxide} + \text{Nitrogen (IV) oxide} + \text{Oxygen}$$

$$\underset{(white)}{2Ca(NO_3)_2(s)} \xrightarrow{heat} \underset{(white)}{2CaO(s)} + \underset{(brown\ gas)}{4NO_2(g)} + O_2(g)$$

$$\text{Zinc nitrate} \xrightarrow{heat} \text{Zinc oxide} + \text{Nitrogen (IV) oxide} + \text{Oxygen}$$

$$\underset{(white)}{2Zn(NO_3)_2(s)} \xrightarrow{heat} \underset{\substack{(yellow\ when\ hot) \\ (white\ when\ cold)}}{2ZnO(s)} + \underset{(brown)}{4NO_2(g)} + O_2(g)$$

$$\text{Lead (II) nitrate} \xrightarrow{heat} \text{Lead (II) oxide} + \text{Nitrogen (IV) oxide} + \text{Oxygen}$$

$$\underset{(white)}{2Pb(NO_3)_2(s)} \xrightarrow{heat} \underset{\substack{(orange\ when\ hot) \\ (yellow\ when\ cold)}}{2PbO(s)} + 4NO_2(g) + O_2(g)$$

$$\text{Copper Nitrate} \xrightarrow{heat} \text{Copper (II) oxide} + \text{Nitrogen (IV) oxide} + \text{Oxygen}$$

$$\underset{(blue)}{2Cu(NO_3)_2(s)} \xrightarrow{heat} \underset{(black)}{2CuO(s)} + \underset{(brown)}{4NO_2(s)} + O_2(g)$$

The general equation for the decomposition of nitrates is shown below;

$$\text{Metal nitrate} \xrightarrow{heat} \text{Metal oxide + Nitrogen (IV) oxide + Oxygen}$$

The nitrates of silver and mercury decompose to give nitrogen (IV) oxide, oxygen and the corresponding metal.

$$\text{Silver nitrate} \xrightarrow{heat} \text{Silver + Nitrogen (IV) Oxide + Oxygen}$$

$$2AgNO_3(s) \xrightarrow{heat} 2Ag(s) + 2NO_2(g) + O_2(g)$$

$$\text{Mercury (II) Nitrate} \xrightarrow{heat} \text{Mercury + Nitrogen(IV) Oxide + Oxygen}$$

$$Hg(NO_3)_2(s) \xrightarrow{heat} Hg(s) + 2NO_2(g) + O_2(g)$$

The general equation for the decomposition of these nitrates is:

$$\text{Metal nitrate} \xrightarrow{heat} \text{Metal + Nitrogen (IV) oxide + Oxygen}$$

The ease with which metal nitrates decompose on heating increases down the reactivity series of metals. Ammonium nitrate decomposes to give steam and nitrogen (I) oxide.

$$\text{Ammonium nitrate} \xrightarrow{heat} \text{Nitrogen (I) oxide + water}$$

$$NH_4NO_3(s) \xrightarrow{heat} N_2O(g) + 2H_2O(l)$$

Effect of heat on sulphates.

Discussion Questions

1. **Which sulphates are not affected by heat? Give a reason.**

   The sulphates of **potassium, sodium, calcium and magnesium** are stable and are not affected by heat.

   However, if hydrated they lose water of crystallisation on heating which condenses high up in the test tube.

$$\text{Hydrated sodium sulphate} \xrightarrow{heat} \text{Sodium sulphate + Water of crystallisation}$$

$$Na_2SO_4.10H_2O(s) \xrightarrow{heat} Na_2SO_4(s) + 10H_2O(l)$$

2. **Write equations for the reactions that occur when some metal sulphates are heated.**

   Zinc sulphate and copper (II) sulphate decompose on strong heating to form the metal oxide, water and sulphur (VI) oxide gas which is colourless. The gas turns moist blue litmus paper red.

$$ZnSO_4.7H_2O(s) \xrightarrow{heat} ZnO(s) + SO_2(g) + 7H_2O(l)$$

(white) *(yellow when hot)*

(white when cold)

$$2CuSO_4.5H_2O(s) \xrightarrow{heat} 2CuO(s) + SO_2(g) + 5H_2O(l)$$

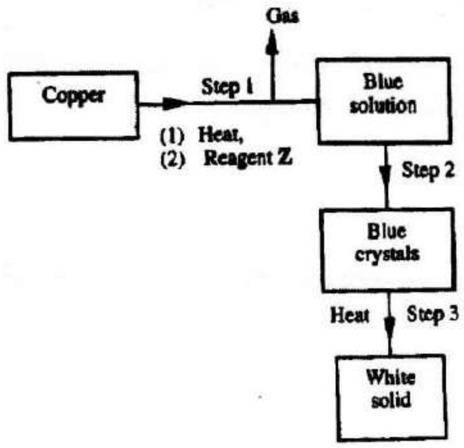
(white) *(black)*

Iron (II) sulphate crystals decompose to give iron (II) oxide, sulphur (IV) oxide, sulphur (VI) oxide and water.

Review Exercises

1. 2006 Q 27 P1

Study the flow chart below and answer the questions that follow.



(a) Name reagent **Z**. (1 mark)
(b) Describe the process which takes place in step 2. (1 mark)
(c) Identify the white solid. (1 mark)

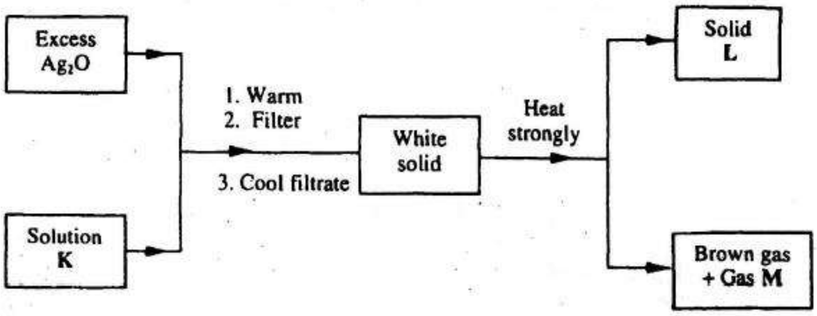
2. 2007 Q 6 P1

In an experiment, a few drops of concentrated nitric acid were added to aqueous iron(II) sulphate in a test-tube. Excess sodium hydroxide solution was then added to the mixture.

(a) State the observations that were made when:
 (i) Concentrated nitric acid was added to aqueous iron (II) sulphate (1 mark)
 (ii) Excess sodium hydroxide was added to the mixture. (1 mark)

(b) Write an ionic equation for the reaction which occurred in (a) (ii) above. (1 mark)

3. 2007 Q 9 P1
Study the flow chart below and answer the questions that follows.

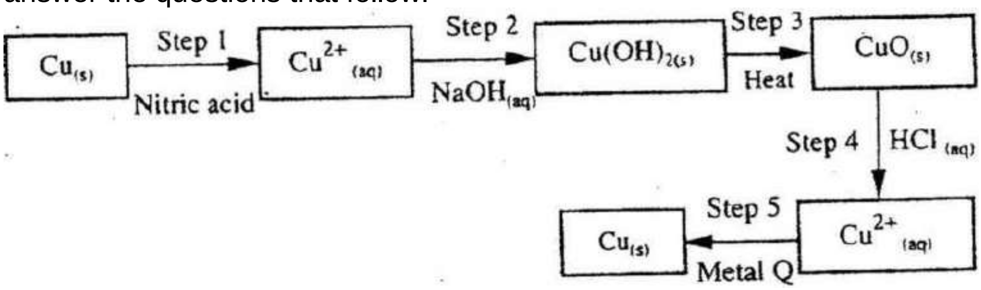


Identify:
(a) Solution K
(b) Solid L
(c) Gas M (3 marks)

4. 2007 Q 13a(i) P1
Name the process that takes place when crystals of zinc nitrate change into solution when exposed to air (1 mark)

5. 2007 Q 18 P1
Starting with sodium metal, describe how a sample of crystals of sodium hydrogen carbonate may be prepared. (3 marks)

6. 2007 Q 3 P2
The flow chart below shows a sequence of chemical reactions starting with copper study it and answer the questions that follow.



(a) In step 1, excess 3M nitric acid was added to 0.5g of copper powder
(i) State two observations which were made when the reactions was in progress. (1 mark)
(ii) Explain why dilute hydrochloric acid cannot be used in step 1 (2 marks)
(iii) I. Write the equation for the reaction that took place in step 1 (1 mark)
II. Calculate the volume of 3M nitric that was needed to react completely with 0.5g of copper powder. (Cu = 63.5) (3 marks)

(b) Give the names of the types of reactions that took place in steps 4 and 5 (1 mark)

(c) Apart from the good conductivity of electricity, state two other properties that make it possible for copper to be extensively used in the electrical industry. (2 marks)

7. 2008 Q 15 P1

The table below gives the solubilities of substances J, K and L at different temperatures

Substance	Solubility in grams per 100 g water at:			
	0°C	20°C	40°C	60°C
J	0.334	0.16	0.097	0.0058
K	27.60	34.0	40.0	45.5
L	35.70	36.0	40.0	37.3

Select the substance which, when dissolved in water, heat is given out. Give a reason. (2 marks)

8. 2008 Q 16

Starting with copper metal, describe how a sample of crystals of copper (II) chloride may be prepared in the laboratory. (3 marks)

9. 2008 Q 17

A compound whose general formula is $M(OH)_3$ reacts as shown by the equation below.

$$M(OH)_{3(s)} + OH^-_{(aq)} \longrightarrow M(OH)_4{}^-_{(aq)}$$
$$M(OH)_{3(s)} + 3H^+_{(aq)} \longrightarrow M^{3+}_{(aq)} + 3H_2O_{(l)}$$

(a) What name is given to compounds which behave like $M(OH)_3$ in the two reactions. (1 mark)

(b) Name two elements whose hydroxides behave like that of M (2 marks

10. 2008 Q 17 P1

When solid B_1 was heated, a gas which formed a white precipitate when passed through lime water was produced. The residue was dissolved in dilute nitric (V) acid to form a colourless solution B_2. when dilute hydrochloric acid was added to solution B_2 a white precipitate which dissolved on warning was formed.

(a) Write the formula of the;

I. Cation in solid B_1 (1 mark)

II. Anion in solid B_1 (1 mark)

(b) Write an ionic equation for the reaction between the residue and dilute nitric (V) acid (1 mark)

11. 2008 Q 2 P2, 2016 Q7 P2

(a) Write an equation to show the effect of heat on the nitrate of:

(i) Potassium (ii) silver (2 marks)

(b) The table below gives information about elements A_1 , A_2 , A_3, and A_4

Element	**Atomic Number**	**Atomic Radius (nm)**	**Ionic radius (nm)**
A_1	3	0.134	0.074
A_2	5	0.090	0.012
A_3	13	0.143	0.050
A_4	17	0.099	0.181

(i) In which period of the periodic table is element A_2? Give a reason.

(ii) Explain why the atomic radius of:

I A_1 is greater than that of A_2;

II A_4 is smaller than its ionic radius (2 marks)

(iii) Select the element which s in the same group as A3 (1 mark)

(iv) Using dots (.) and crosses(x) to represent outermost electrons. Draw a diagram to show the bonding in the compound formed when A_1 reacts with A_4 (1 mark)

12. 2009 Q 11 P1

Starting with 50 cm^3 of 2.8M sodium hydroxide, describe how a sample of pure sodium sulphate crystals can be prepared. (3 marks)

13. 2009 Q 18 P1

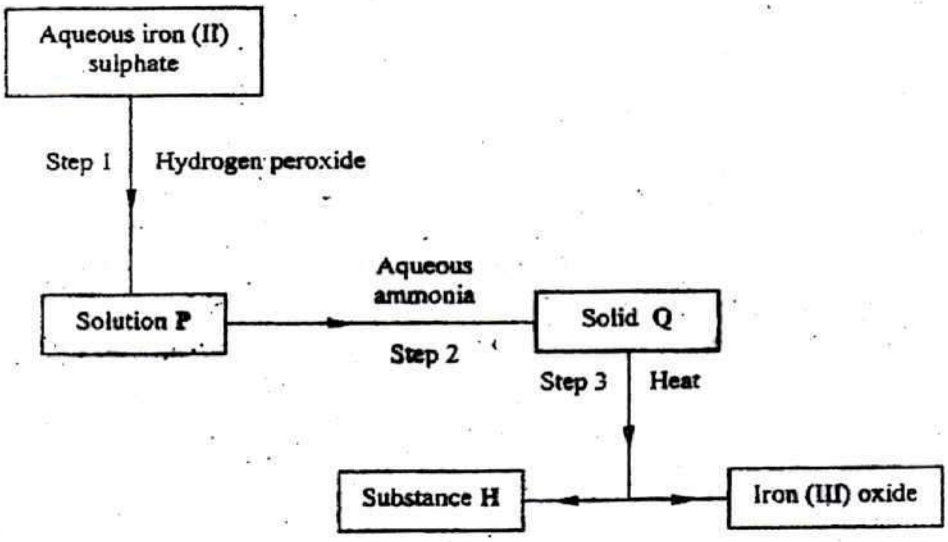
Bottle of sodium carbonate, sodium chloride and sugar have lost their labels. A Student prepares and tests an aqueous solution of a sample from each bottle. The results obtained are as shown in the table below.

Bottle	PH	Electrical conductivity	Correct label
1	7	conducts	
2	7	Does not conducts	
3	10	conducts	

Complete the table by filling the correct label for each bottle. (3 marks)

14. 2009 Q 23 P1

Use the flow chart below to answer the questions that follow.



(a) What observation would be made in step I (1 mark)
(b) Name another substance that could be used in step 2 (1 mark)
(c) Give the name of substance H. (1 mark)

15. 2009 Q 25 P1

For each of the following reactions, state the observation and write the formula of the compound responsible for the observation

(a) Bromide water is added to aqueous potassium iodine (1½ marks)

(b) Excess aqueous ammonia is added to copper (II) hydroxide (precipitate) (1½ marks)

16. 2010 Q 1 P1

(a) Distinguish between a deliquescent and a hygroscopic substance. (2 marks)

(b) Give one use of hygroscopic substances in the laboratory. (1 mark)

17. 2010 Q 5 P1

Hydrated cobalt (II) chloride exists as pink crystals and anhydrous cobalt (II) chloride is a blue powder. Describe a laboratory experiment that can be used to show that the action of heat on hydrated cobalt (II) chloride is a reversible reaction.
(3 marks)

18. 2010 Q 24 P1

Describe how a solid sample of the double salt, ammonium iron (II) sulphate, can be prepared using the following reagents; Aqueous ammonia, sulphuric (VI) acid and iron metal.
(3 marks)

19. 2011 Q 3 P1

A mixture contains ammonium chloride, copper (II) oxide and sodium chloride. Describe how each of the substances can be obtained from the mixture (3 marks)

20. 2011 Q 13 P1

Distinguish between the terms deliquescent and efflorescent as used in chemistry.
(2 marks)

21. 2011 Q 30 P1

A sample of river water is suspected to contain zinc ions. Describe how the presence of zinc ions and sulphate ions can be established. (3 marks)

22. 2011 Q 4 P2

(a) When excess calcium metal was added to 50 cm^3 of 2 M aqueous copper (II) nitrate in a beaker, a brown solid and bubbles of gas were observed.

(i) Write two equations for the reactions which occurred in the beaker. (2 marks)

(ii) Explain why it is not advisable to use sodium metal for this reaction. (2 marks)

(b) Calculate the mass of calcium metal which reacted with copper (II) nitrate solution. (relative atomic mass of Ca=40) (2 marks)

(c) The resulting mixture in (a) above was filtered and aqueous sodium hydroxide added to the filtrate drop wise until in excess. What observations were made? (1 mark)

(d) (i) Starting with calcium oxide, describe how a sample of calcium carbonate can be prepared. (3 marks)

(ii) Name one use of calcium carbonate. (1 mark)

23. 2012 Q6 P1

Study the information in the table below and answer the questions that follow:

Salt	**Solubility (g/100g water)**	
	At 40°C	At 60°C
CuSo4	28	38
Pb(NO3)2	79	98

A mixture containing 35g of CuSO4 and 78g of Pb(NO3)2 in 100g of water at 60°C was cooled to 400C.

(a) Which salt crystallised out? Give a reason (2 marks)

(b) Calculate the mass of the salt that crystallised out. (1 mark)

24. 2012 Q16 P1

Use the following information on substances S, T, V and hydrogen to answer the questions that follow:

(i) T displaces V from a solution containing V ions

(ii) Hydrogen reacts with the heated oxide of S but has no effects on heated oxide of V.

(a) Arrange substances S, T, V and hydrogen in the order of increasing reactivity.

(1 mark)

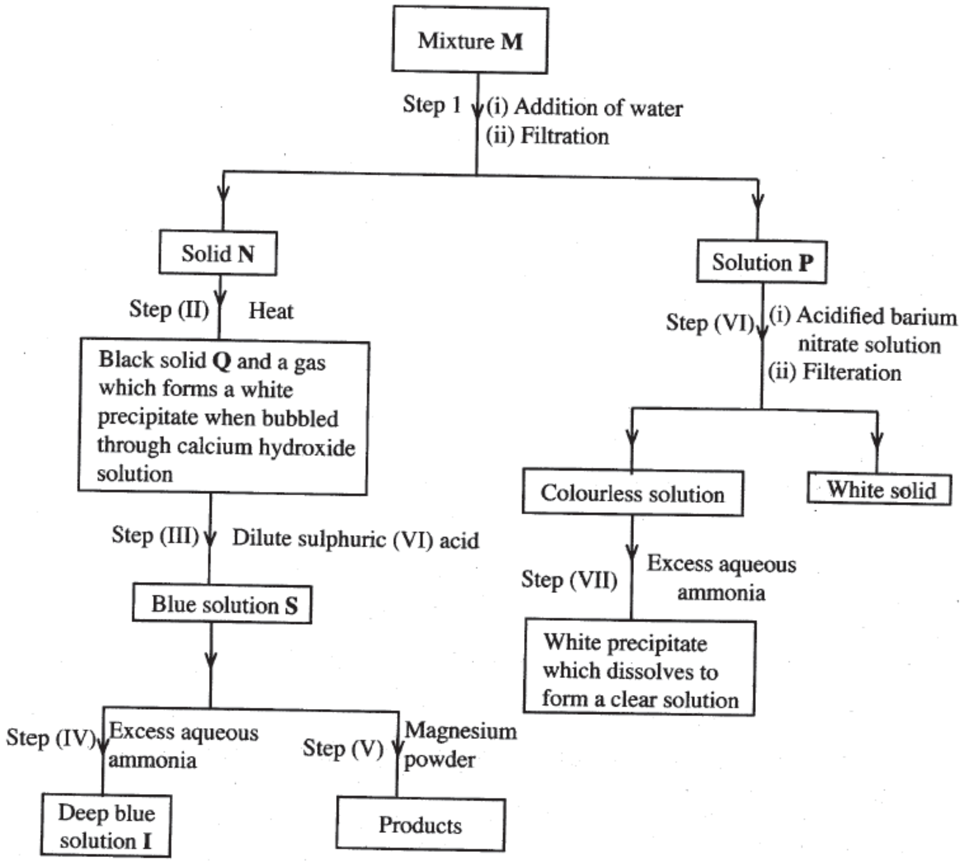
(b) If T and V are divalent metals, write an ionic equation for the reaction in (i) above. (2 marks)

25. 2012 Q25 P1

Describe how a solid sample of potassium sulphate can be prepared starting with 200cm^3 of 2M potassium hydroxide. (3 marks)

26. 2012 Q6 P2

The flow chart below shows a sequence of reaction involving a mixture of two salts, mixture M. study it and answer the questions that follow.



(a) Write the formula of the following
(i) Anion in solid **Q** (1 mark)
(ii) The two salts present in mixture **M** (2 marks)

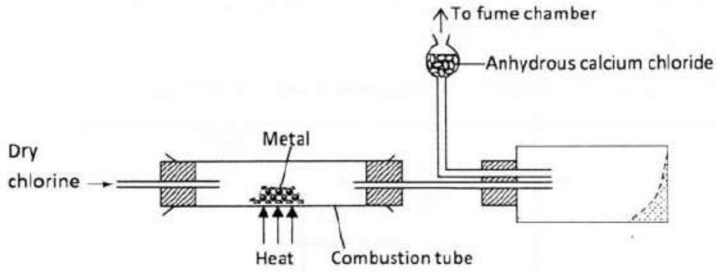
(b) Write an ionic equation for the reaction in step (VI) (1 mark)

(c) State and explain the observations made in step (V) (3 marks)

(d) (i) Starting with Lead (II) oxide, describe how a pure solid sample of lead sulphate can be prepared in the laboratory (2 marks)
(ii) How can one determine whether the led sulphate prepared is pure? (2 marks)

27. 2013 Q2 P1
Write equations to show the effect of heat on each of the following:
(a) Sodium hydrogen carbonate (1 mark)
(b) Silver nitrate (1 mark)
(c) Anhydrous iron (II) sulphate (1 mark)

28. 2013 Q5 P1
The diagram below illustrates a method of preparing salts by direct synthesis



(a) This method can be used to prepare either aluminium chloride or iron (III) chloride. Explain why it cannot be used to prepare sodium chloride. (2 marks)

(b) Describe how a sample of sodium chloride can be prepared in the laboratory by direct synthesis. (2 marks)

29. 2013 Q26 P1

By using aqueous sodium chloride, describe how a student can distinguish calcium ions from lead ions.

30. 2013 Q5a-b P2

(a) Describe one method that can be used to distinguish between sodium sulphate and sodium hydrogen sulphate. (2 marks)

(b) Describe how a pure sample of lead (II) sulphate can be prepared in the laboratory starting with lead metal. (2 marks)

31. 2014 Q 2 P1

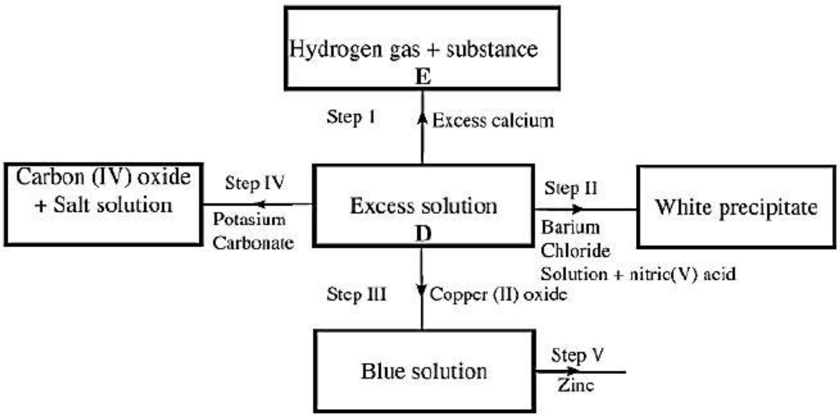
When dilute hydrochloric acid was reacted with solid B, a colourless gas which extinguished a burning splint was produced. When an aqueous solution of solid B was tested with a blue litmus paper, the paper turned red / pink.

(a) Identify the anion present in solid B. (1 mark)

(b) Write an ionic equation for the reaction between solid B and dilute hydrochloric acid. (1 mark)

32. 2014 Q5 P2

(a) The scheme below shows some of the reaction of solution D. Study it and answer the questions that follow



(i) Give a possible caution present in solution D (1 mark)

(ii) Write an ionic equation for the reaction in Step II (1 mark)

(iii) What observations would be made in Step V? Give a reason (2 marks)

(iv) Explain why the total volume of hydrogen gas produced in step 1 was found to be very low although calcium and solution D were in excess. (2 marks)

(v) State one use of substance E. (1 mark)

(b) Starting with solid sodium chloride, describe how a pure sample of lead (II) Chloride can be prepared in the laboratory (3 marks)

(c) (i) State a property of anhydrous calcium chloride which makes it suitable for use as a drying agent for chlorine gas. (1 mark)
(ii) Name another substance that can be used to dry chlorine gas (1 mark)

33. 2015 Q8 P1
When solid A was heated strongly, it gave off water and a solid residue. When water was added to the solid residue, the original solid A, was formed
(a) What name is given to the process described? (1 mark)
(b) Give one example of solid A (1 mark)

34. 2015 Q21 P1
Describe how samples of lead (II) sulphate, ammonium chloride and sodium chloride can be obtained from a mixture of the three. (3 marks)

35. 2015 Q25 P1
Starting with barium nitrate solution, describe how a pure sample of barium carbonate can be prepared in the laboratory. (3 marks)

36. 2016 Q5 P1
Starting with sodium metal, describe how a sample of crystals of sodium hydrogen carbonate may be prepared. (3 marks)

37. 2017 Q15 P1.
Starting with copper, describe how a pure sample of copper (II) carbonate can be prepared. (3 marks)

38. 2018 Q15 P1.
You are provided with solid potassium hydrogen carbonate. Describe how a solid sample of potassium nitrate can be prepared. (3 marks)

39. 2019 P1 Q24.
Starting with copper turnings, describe how a sample of copper (II) Sulphate crystals can be prepared in the laboratory. (2 marks)