Chapter 18
SALTS

Arrangement of Ions | Melting Points | Conduction of Electricity
Soluble & Insoluble Salts | Preparation of Soluble Salts

DESCRIPTIVE QUESTIONS – Arrangement of Ions in Salts

Q.1 Define salts. How are the oppositely charged ions of sodium chloride arranged in a crystal?

Salt: An ionic compound formed by the neutralization reaction of an acid and a base, along with water, is called a salt. *Example:* Sodium chloride, potassium iodide, barium chloride.

When acids neutralize bases (alkalis), salts and water are formed. This is called a **neutralization reaction**.

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

Arrangement of positive and negative ions of a salt:

- A salt is a chemical combination of positive and negative ions, bonded together by electrostatic force of attraction.
- The strength of attraction depends on the magnitude of the charges and the distance between the ions.
- These oppositely charged ions arrange themselves in a regular, repeating pattern forming a three-dimensional network called a **crystal lattice**.
- Under ordinary conditions of temperature and pressure, ionic compounds exist as **crystalline solids**.

Crystal lattice of sodium chloride: In NaCl, each sodium ion is surrounded by six chloride ions, and each chloride ion is surrounded by six sodium ions, arranged as a **face-centred cube**. The bigger chloride ions occupy the corners and the centre of each face of the cube, while the smaller sodium ions occupy the edges of the cube.

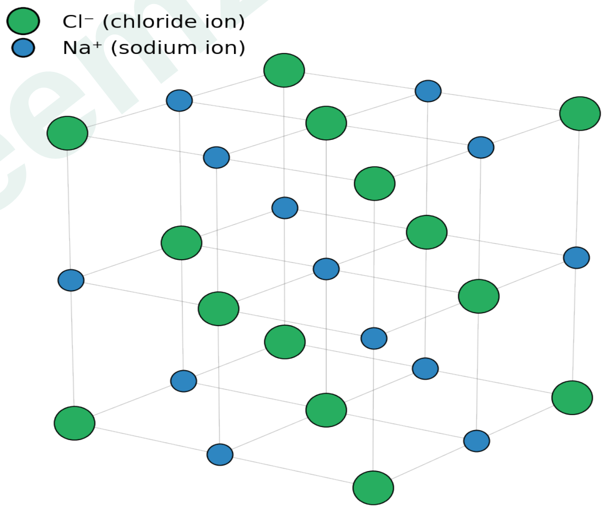

Fig. 18.1 – Crystal lattice of sodium chloride (face-centred cubic arrangement)

Interesting Information
How was common salt used in the past? It was a valuable commodity – even used as currency.

MELTING POINTS OF IONIC COMPOUNDS

Q.2 Why are melting points of ionic salts generally very high?

As very strong electrostatic forces of attraction exist between the ions within a crystal lattice, a significant amount of energy is needed to break these forces – this is why melting points of ionic compounds are generally very high. During melting, these attractive forces break down and let the ions move freely.

Factors affecting melting points: Melting points depend upon the charges present on the ions and their sizes. Higher charges and smaller ion sizes lead to stronger attractions and hence higher melting points.

CONDUCTION OF ELECTRICITY BY IONIC COMPOUNDS

Q.3 Why do aqueous solutions of ionic compounds serve as strong electrolytes?

- Ionic compounds do **not** conduct electricity in the solid state – ions are held together by strong forces and are not free to move.
- **Molten state:** when heated to their melting point, ions become free to move and hence conduct electricity.
- **Aqueous solution:** when dissolved in water, forces of attraction between ions break down and ions again become free to move.
- Ionic compounds serve as **strong electrolytes** because they dissociate completely into ions when dissolved in water.

18.1 Quick Check
1. How do the ions in NaCl arrange to form a crystal lattice?
Ans. Each sodium ion is surrounded by six chloride ions and each chloride ion by six sodium ions, forming a face-centred cube; chloride ions sit at the corners and face-centres, sodium ions occupy the edges.

2. Do you expect the melting point of KCl to be higher or lower than NaCl?
Ans. NaCl has a slightly higher melting point (801°C) than KCl (770°C). Na^+ is smaller than K^+, so it forms a stronger electrostatic attraction with Cl^-, needing more energy to break – giving NaCl the higher melting point.

SOLUBLE AND INSOLUBLE SALTS

Q.4 Describe which salts are soluble in water and which are insoluble.

Several factors affect the solubility of a salt, including the **nature of the salt** and the **temperature**.

- **Soluble salts:** all sodium, potassium and ammonium salts; all metallic nitrates; most chlorides and sulfates.
- **Insoluble salts:** all carbonates except those of sodium, potassium and ammonium; chlorides of silver and lead; sulfates of barium and lead.
- **Hydroxides:** sodium, potassium and ammonium hydroxides are soluble; calcium hydroxide is only partially soluble; all other hydroxides are insoluble.

Table 18.1 – Solubility of salts in water

Salts	Soluble	Insoluble
Sodium, potassium, ammonium & all nitrates	All	None
Chlorides, bromides, iodides	Mostly soluble	Silver, Lead (II)
Sulphates	Mostly soluble	Barium, Lead (II), Calcium
Carbonates	Sodium, Potassium, Ammonium	Mostly insoluble
Hydroxides	Sodium, Potassium, Ammonium, Calcium (partially)	Mostly insoluble

PREPARATION OF SOLUBLE SALTS

Q.5 Explain the methods of preparation of soluble salts.

(a) Reaction of water-soluble acid with water-soluble base – Titration method: involves the reaction of a water-soluble acid and a water-soluble base, selected depending on the salt to be prepared.

Preparation of potassium nitrate:

- Appropriate volumes of potassium hydroxide and nitric acid are reacted together in a conical flask.

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

- **Evaporation:** transfer the solution to an evaporating dish; evaporate to dryness (if stable to heat) or until a thin film of crystals forms.
- **Cooling:** cool the solution slowly, then filter to get pure crystals of the salt.
- **Drying:** dry the crystals between folds of filter paper; a desiccator or oven can also be used.

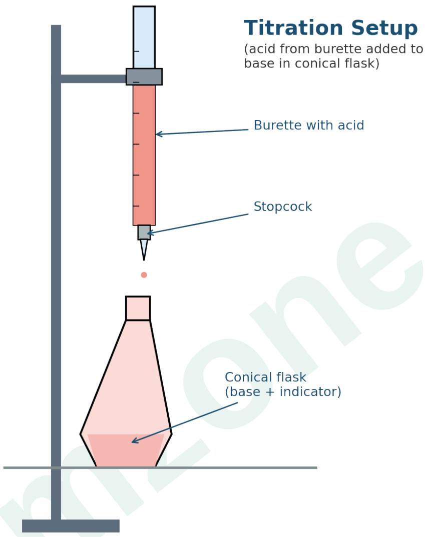

Fig. 18.2a – Titration setup used to prepare a soluble salt

(b) Reaction of an acid with an insoluble base – Neutralization reaction: the insoluble base may be a metal oxide, metal carbonate or metal hydroxide; a salt is also formed when an acid reacts with a metal.

- Mix the selected acid with the selected base/metal with constant stirring until the acid has fully reacted.
- The insoluble base or metal is always added in **excess** to ensure all the acid has reacted.
- Filter the mixture to remove the excess base/metal.
- Evaporate the filtrate to get a saturated solution.
- Allow the solution to cool so pure crystals start appearing.
- Filter and dry the crystals carefully.

Crystallisation in an Evaporating Dish

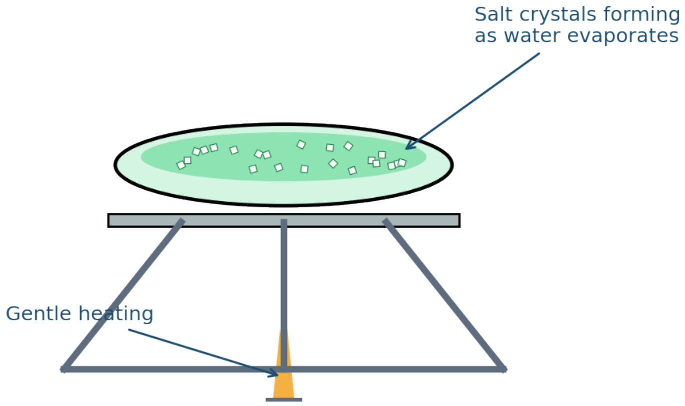

Fig. 18.2b – Evaporating dish: crystallisation by gentle heating

Reactions involved:

$$Cu(OH)_2(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + 2H_2O(l)$$

$$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$$

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

Q.6 Write a method to prepare pure crystals of sodium sulphate.

- Add 50 cm^3 of 1 mol/dm^3 sodium hydroxide solution to a clean conical flask using a pipette.
- Add a few drops of phenolphthalein indicator – solution turns pink.
- Run 1 mol/dm^3 sulphuric acid from the burette until the colour changes from light pink to colourless.
- Note the volume of acid used to neutralize the sodium hydroxide.

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

- Repeat using the same volumes of NaOH and acid, but without the indicator.
- Transfer the neutralized solution to an evaporating dish and heat gently until one-third remains.
- Dip a glass rod into the hot solution – if the immersed end turns cloudy, it is concentrated enough for crystals to form.
- Cool the solution for complete crystallisation.
- Filter and wash crystals with a little cold distilled water to remove soluble impurities.
- Dry the crystals within filter paper folds or in an oven.

Q.7 How will you prepare pure crystals of zinc sulphate?

- Take about 25 cm^3 dilute sulphuric acid in a beaker and warm it gently.
- Add small granules of zinc metal slowly with constant stirring until metal starts settling at the bottom.
- Filter to remove undissolved zinc and collect the filtrate.
- Evaporate the filtrate in an evaporating dish by gentle heating to concentrate the solution.
- Cool the solution slowly and allow crystals to appear.
- Filter to collect the crystals, then dry on filter paper or in an oven.

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

18.2 Quick Check
1. Which salts of barium and calcium are soluble in water?
Ans. Barium nitrate, barium chloride and barium hydroxide are soluble; calcium nitrate and calcium chloride are water-soluble salts of calcium.

2. How is barium sulphate prepared in the laboratory?
Ans. By precipitation between a soluble barium salt and dilute sulfuric acid:

$$BaCl_2(aq) + H_2SO_4(aq) \rightarrow BaSO_4(s) + 2HCl(aq)$$

A white precipitate of $BaSO_4$ forms because it is insoluble in water.

MULTIPLE CHOICE QUESTIONS (EXERCISE)

1. Which base is not soluble in water?
(a) KOH (b) $Mg(OH)_2$ (c) NaOH (d) Na_2CO_3

2. The shape of the crystal of NaCl is:
(a) Cubic (b) Hexagonal (c) Rombic (d) Trigonal

3. Which salts are always soluble in water?
(a) Chlorides (b) Sulphates (c) Nitrates (d) Carbonates

4. Which salt forms when marble reacts with dilute nitric acid?
(a) Sodium nitrate (b) Calcium nitrate (c) Potassium nitrate (d) Magnesium nitrate

5. Water insoluble salt is:
(a) Calcium sulphate (b) Sodium sulphate (c) Potassium sulphate (d) Magnesium sulphate

6. Oppositely charged ions in salt are bonded together by a force called:
(a) Centripetal force (b) Ionic force (c) Dipole–Dipole force (d) Electrostatic force

7. The strength of attraction between charges depends upon:
(a) Magnitude of the charges (b) Distance between ions (c) Both a & b (d) None of a & b

8. Three-dimensional network structure of oppositely charged ions is called:
(a) A cell (b) Crystal Lattice (c) Cubic structure (d) None of all

9. In NaCl, each sodium ion is surrounded by how many chloride ions?
(a) 4 (b) 5 (c) 6 (d) 7

10. The shape of sodium chloride crystal is:
(a) Hexagonal (b) Face-centered cube (c) Spherical (d) Pentagonal

11. Which ions are present at the corners of each face of the cube in NaCl?
(a) Sodium ions (b) Chloride ions (c) Both a & b (d) None of these

12. Where are sodium ions located in the NaCl cube?
(a) At corners (b) At the centre (c) At edges (d) Everywhere in cube

13. Which salt was used as currency in the past?
(a) Silver chloride (b) Sodium chloride (c) Barium chloride (d) All of above

20. Which property of ionic compounds contributes to their use as a strong electrolyte?
(a) Having crystal lattice (b) Having high melting point (c) Solid at room temperature (d) Dissociate completely into ions

21. Salts of which ions are mostly soluble:
(a) Sodium (b) Potassium (c) Ammonium (d) All of these

22. Most metallic nitrates are:
(a) Soluble in water (b) Insoluble in water (c) Partially soluble (d) Soluble on heating

23. All carbonates are insoluble in water except of:
(a) Sodium (b) Potassium (c) Ammonium (d) All of these

24. Chlorides of which ion are regarded as insoluble:
(a) Sodium (b) Potassium (c) Silver (d) Magnesium

25. Which sulphate is insoluble in water?
(a) Barium sulfate (b) Potassium sulfate (c) Ammonium sulfate (d) Sodium sulfate

26. Which hydroxide is only partially soluble in water?
(a) Sodium hydroxide (b) Potassium hydroxide (c) Calcium hydroxide (d) Aluminium hydroxide

27. When acid and base react together forming salt & water, this reaction is termed:
(a) Displacement reaction (b) Neutralization reaction (c) Oxidation–reduction reaction (d) Endothermic reaction

28. Reaction of water-soluble acid with water-soluble base involves a method called:
(a) Crystallization (b) Evaporation (c) Titration (d) Condensation

29. Which base reacts with nitric acid to produce potassium nitrate and water?
(a) Potassium hydroxide (b) Potassium carbonate (c) Sodium hydroxide (d) Sodium carbonate

30. Copper sulphate is used as:
(a) Flavoring agent (b) Cleaning agent (c) Water softening agent (d) Fungicide

31. Evaporating dish is commonly called as:
(a) Funnel (b) Filtering dish (c) China dish (d) Condensation chamber

32. Sodium sulphate is prepared in the laboratory by adding sodium hydroxide and:
(a) Barium chloride (b) Hydrochloric acid (c) Ammonium sulphate (d) Sulphuric acid

14. Melting points of ionic compounds are generally:
(a) Very high (b) Very low (c) Moderate (d) Less than water

15. During which process do the attractive forces break down and ions move freely?
(a) Melting (b) Boiling (c) Condensing (d) Evaporating

16. Melting point of ionic compounds depend upon their:
(a) Charges on the ions (b) Sizes of the ions (c) Both a & b (d) Physical state

17. Compare the melting points of KCl and NaCl.
(a) Both are equal (b) M.P. of NaCl is lower than KCl (c) M.P. of NaCl is very low (d) M.P. of NaCl is higher than KCl

18. Ionic compounds do not conduct electricity in:
(a) Solid state (b) Molten state (c) Aqueous solution (d) Fused form

19. In solid state, which property is exhibited by the ions?
(a) Held together by strong electrostatic force (b) Not free to move (c) Cannot conduct electricity (d) All of these

33. Colour of phenolphthalein in basic solution is:
(a) Colourless (b) Light pink (c) Yellow (d) Orange

34. Reaction of an acid with an insoluble base is also a:
(a) Neutralization reaction (b) Crystallization reaction (c) Displacement reaction (d) Substitution reaction

35. Which step ensures that the acid has completely reacted with the base during neutralization?
(a) Filtration (b) Evaporation (c) Heating (d) Constant stirring

36. Which base is used for the formation of copper (II) chloride?
(a) Sodium hydroxide (b) Copper (II) hydroxide (c) Magnesium hydroxide (d) Copper (I) carbonate

37. Zinc reacts with dilute sulphuric acid to produce:
(a) Zinc chloride (b) Zinc hydride (c) Zinc hydroxide (d) Zinc sulphate

ANSWER KEY

1	2	3	4	5	6	7	8	9	10
b	a	a	b	a	d	c	b	c	b
11	12	13	14	15	16	17	18	19	20
b	c	b	a	a	c	d	a	d	d
21	22	23	24	25	26	27	28	29	30
d	a	d	c	a	c	b	c	a	d
31	32	33	34	35	36	37			
c	d	b	d	d	b	d			

SHORT ANSWER QUESTIONS (EXERCISE)

Q1. Which factors are responsible for the strength of electrostatic forces between ions?

The strength depends upon: (i) The magnitude of the charges – higher charge, stronger force. (ii) The size of ions – smaller ions, stronger force.

Q2. Is lead (II) chloride soluble in water? Give comments.

Lead (II) chloride is slightly soluble in water because the high lattice energy of its crystals is not fully overcome by the hydration energy released. Solubility is 1.1 g at 25°C, rising to 3.3–3.5 g at 100°C.

Q3. Which lead salts are insoluble in water?

Lead (II) sulphate ($PbSO_4$), lead (II) carbonate ($PbCO_3$), lead (II) phosphate ($Pb_3(PO_4)_2$), lead (II) sulphide (PbS), lead (II) hydroxide ($Pb(OH)_2$) and lead (II) chromate ($PbCrO_4$) are all insoluble. Lead (II) salts are generally insoluble except nitrates, acetates and chlorates.

Q4. Name two insoluble carbonates.

Carbonates of most metals are insoluble except those of sodium, potassium and ammonium. Two examples: calcium carbonate ($CaCO_3$) and lead (II) carbonate ($PbCO_3$).

Q5. What is a crystal lattice?

A crystal lattice is a regular, repeating three-dimensional arrangement of particles (atoms, ions or molecules) in a solid – the ordered pattern that makes up a crystal. *Example:* in NaCl, Na^+ and Cl^- ions are arranged in a repeating face-centred cube pattern.

SLO BASED SHORT ANSWER QUESTIONS

Q6. Define neutralization reaction.

It is a reaction in which acids neutralize bases or alkalis, forming salt and water. *Example:* sodium hydroxide neutralizes hydrochloric acid to form sodium chloride and water.

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

Q7. What is a salt? Give an example.

An ionic compound formed by the neutralization reaction of an acid and a base, along with water, is called a salt. *Example:* sodium chloride, potassium iodide, barium chloride.

Q8. Draw the crystal lattice of sodium chloride.

See Fig. 18.1 above – face-centred cube of Na^+ and Cl^- ions.

Q9. What are the factors affecting melting points of ionic compounds?

They depend upon the charges present on the ions and their sizes: higher charges and smaller ion sizes give stronger attractions and hence higher melting points.

Q10. Why do ionic compounds not conduct electricity in the solid state?

Because there are no free ions – in the solid state, ions are held together by strong forces of attraction and are not free to move.

Q11. Why do ionic compounds conduct electricity in aqueous solution and in the molten state?

When heated to their melting point, or dissolved in water, the forces of attraction between ions break down and the ions become free to move – allowing them to conduct electricity.

Q12. Enlist some soluble chlorides.

All chlorides of metallic ions except silver and lead (II) chlorides are soluble in water. *Examples:* NaCl, KCl, $MgCl_2$.

Q13. Which carbonates are soluble in water?

Carbonates of sodium, potassium and ammonium are soluble in water; others are insoluble.

Q14. Which salts are generally soluble?

All sodium, potassium and ammonium salts are soluble, e.g. NaCl, Na_2SO_4, K_2CO_3. All nitrate salts are generally soluble, e.g. KNO_3, $Ca(NO_3)_2$.

Q15. Write the reactions for the manufacturing of sodium sulphate and zinc sulphate.

$$\text{(i)}\ 2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

$$\text{(ii)}\ Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$

Q16. How will you prepare copper (II) chloride, magnesium sulphate and calcium chloride using dilute acids?

$$Cu(OH)_2(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + 2H_2O(l) \quad \text{(copper (II) chloride)}$$

$$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g) \quad \text{(magnesium sulphate)}$$

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l) \quad \text{(calcium chloride)}$$

Q17. How can you prepare a soluble salt using an excess of insoluble base?

A soluble salt can be prepared using an excess of an insoluble base reacted with an acid like sulphuric acid. *Example:* copper (II) sulphate is prepared by adding insoluble CuO to sulphuric acid.

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

Q18. What role does water play in dissolving salts?

Water molecules, being polar, surround and separate the ions in the salt, reducing the electrostatic attraction between them. This allows the ions to move freely, get dissolved, and participate in conduction of electricity.

Q19. What are the causes of hardness of ionic compounds?

Ionic compounds are hard because strong ionic bonds form a stiff lattice structure. Each ion is tightly surrounded by oppositely charged ions, resisting deformation and giving the material its hardness.

Q20. How can soluble salts be prepared using excess metal?

Soluble salts can be prepared by reacting an acid with an excess amount of metal; the metal reacts with the acid to produce a salt and hydrogen gas. The excess metal ensures complete reaction.

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

Q21. What is the purpose of titration in the preparation of sodium chloride?

- We can measure the exact volume of acid and base used to neutralize the given volume of the other reacting substance.
- Titration also helps to determine the concentration of the solution.

Q22. What indicator is used in the strong acid–base titration method, and why?

Phenolphthalein is used because it changes colour from colourless to light pink at the end point of the reaction, indicating that the strong acid and strong alkali have fully reacted.

Q23. Why is excess zinc added during the preparation of zinc sulphate?

Excess zinc ensures that all the acid has reacted completely. When no more bubbles are observed, it indicates the acid is fully consumed.

Q24. Which gas is evolved when calcium carbonate reacts with hydrochloric acid?

Carbon dioxide gas (CO_2) is evolved, observed as effervescence (bubbling).

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

Q25. How is calcium chloride obtained after reacting $CaCO_3$ with HCl?

After $CaCO_3$ and HCl are reacted and effervescence stops, the mixture is filtered to remove excess $CaCO_3$. The filtrate, containing calcium chloride, is evaporated and cooled to obtain calcium chloride crystals.

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

CONSTRUCTED RESPONSE QUESTIONS

1. Why are melting points of ionic salts generally very high?

The melting points of ionic salts are very high because of the strong electrostatic forces of attraction between oppositely charged ions.

- Ionic salts like NaCl and MgO consist of positive and negative ions.
- These ions are arranged in a regular, repeating pattern to give three-dimensional network structures called crystal lattices.
- Each ion is surrounded by a number of oppositely charged ions.
- To melt a salt, a large amount of energy is needed to overcome these strong forces.

Example: melting point of sodium chloride is 801°C and that of MgO is 2852°C.

2. Why do ionic salts exist in the solid state?

Ionic salts exist in the solid state at room temperature because of the strong electrostatic forces of attraction between their oppositely charged ions.

- Ionic salts are made of positive (cations) and negative (anions) ions.
- Oppositely charged ions attract each other strongly and arrange into a giant crystal lattice, with forces holding the ions together strongly in all directions.

Sodium chloride has a higher melting point than magnesium chloride. Although Mg^{2+} has a higher charge (+2) than Na^+, $MgCl_2$ shows some covalent character because:

- Mg^{2+} is small and highly charged.
- It strongly polarizes the Cl^- ion.
- This reduces purely ionic bonding strength.

As a result, $MgCl_2$ melts at a slightly lower temperature than NaCl.

3. How will you prepare calcium sulphate in the laboratory?

Preparation of calcium sulphate (insoluble salt): calcium sulphate is prepared by reacting calcium carbonate with dilute sulphuric acid.

$$CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + CO_2(g) + H_2O(l)$$

- Take calcium carbonate (marble chips) in a beaker or flask.
- Add dilute sulphuric acid slowly.
- Effervescence occurs due to evolution of CO_2 gas.
- Calcium sulphate forms as a white solid and, being insoluble in water, precipitates out.

4. Compare the melting points of NaCl and $MgCl_2$.

Melting point of NaCl is 801°C; melting point of $MgCl_2$ is 714°C. Sodium chloride has a higher melting point than magnesium chloride, for the reasons explained in Q.2 above.

5. How would you prepare pure crystals of copper sulphate in the laboratory?

Pure crystals of copper sulphate can be prepared from sulphuric acid (H_2SO_4) and copper (II) oxide (CuO).

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

- Measure 25 cm^3 dilute sulphuric acid into a beaker.
- Add excess copper (II) oxide to the dilute acid.
- The reaction produces copper (II) sulphate and water; continue adding CuO until all the acid has reacted.
- **Separation:** filter the mixture to remove unreacted CuO.
- **Crystallisation:** evaporate the filtrate to obtain a hot saturated solution; allow it to cool and crystallize, then filter to collect the crystals; dry in a desiccator or oven.

Result: pure crystals of copper sulphate ($CuSO_4$) are obtained.

DESCRIPTIVE QUESTIONS (EXERCISE)

1. Explain the formation of sodium chloride crystal.

See Q.1 from theory section above.

2. Explain the methods of preparation of two soluble salts.

See Q.5 from theory section above.

3. How are the following salts prepared? K_2SO_4, $CuCl_2$, $CaCl_2$

(i) Preparation of K_2SO_4

$$2KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$$

- Take 50 cm^3 of KOH solution in a beaker or conical flask with a pipette.
- Add a few drops of phenolphthalein indicator to get a pink colour.
- Add dilute sulphuric acid slowly with constant stirring until the solution becomes colourless (neutral).
- Note the volume of acid used to neutralize the potassium hydroxide.
- Repeat the experiment with the same volumes of the two reactants, without using the indicator.
- Heat the neutral solution in an evaporating dish gently until one-third volume is left.
- Cool down the solution for crystallization to occur completely.
- Filter and dry the crystals of potassium sulphate.

Result: pure crystals of potassium sulphate (K_2SO_4) are prepared.

(ii) Preparation of copper (II) chloride ($CuCl_2$)

This method involves a neutralization reaction between the soluble acid and an insoluble base like copper hydroxide and dilute hydrochloric acid.

$$Cu(OH)_2(s) + 2HCl(aq) \rightarrow CuCl_2(s) + 2H_2O(l)$$

- Take a small amount of copper (II) hydroxide in a beaker.
- Add dilute hydrochloric acid slowly while stirring.
- The blue solution gradually diminishes, forming a green solution of copper (II) chloride.
- Continue adding acid until the base is fully utilized, forming salt.
- Filter if any impurity remains.
- Evaporate the filtrate to get a saturated solution.
- Allow it to cool to obtain green crystals of copper (II) chloride.
- Filter and dry the crystals carefully.

Result: copper (II) chloride is prepared in the form of green hydrated crystals.

(iii) Preparation of calcium chloride ($CaCl_2$)

Calcium chloride is prepared from hydrochloric acid (HCl) and calcium carbonate ($CaCO_3$).

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

- Measure 25 cm^3 dilute hydrochloric acid into a beaker.
- Add excess calcium carbonate to the dilute HCl in the beaker.
- The reaction produces calcium chloride, carbon dioxide and water.
- Carbon dioxide evolves with effervescence, indicating that all the acid has reacted.
- Filter the mixture to remove unreacted $CaCO_3$.
- Evaporate the filtrate to obtain a hot saturated solution. Allow the solution to cool and crystallize.

Result: pure crystals of calcium chloride ($CaCl_2$) are prepared.

4. Describe which salts are soluble in water and which are insoluble.

See Q.4 from theory section above.