

## SYLLABUS

- (i) Water as a universal solvent:
  - Solutions as 'mixtures' of solids in water; saturated solutions.
  - Qualitative effect of temperature on solubility (e.g. solutions of calcium sulphate, potassium nitrate, sodium chloride in water).
- (ii) Hydrated and **anhydrous** substances
  - (a) Hydrated substances  
*Water of crystallisation – meaning and examples*
  - (b) **Anhydrous** substances :  
*Meaning and examples only*
  - (c) Properties :
    - Efflorescence • Deliquescence • Hygroscopy  
*(Definition and examples of each of the above)*
- (iii) Drying and dehydrating agents  
*Meaning and examples only*
- (iv) Soft water and hard water.
  - Meaning (in terms of action of soap)
  - Advantages and disadvantages of soft water and hard water.
  - Types and causes of hardness.
  - Removal of hardness (i) by boiling (ii) by addition of washing soda.

## 3 A. WATER

**Formula** :  $\text{H}_2\text{O}$

**Chemical name** : Dihydrogen oxide

**Molecular mass** :  $2 \times 1 + 16 \times 1 = 18$  amu

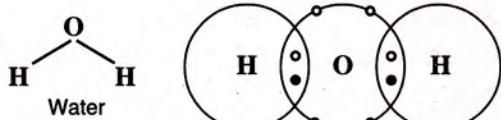


Fig. 3.1 Covalent bonding in a water molecule

### 3.1 INTRODUCTION

Water is the most important natural resource; more than 70% of Earth's area is occupied with water. However, **hardly 2·5% of this water makes up the world's supply of fresh water, including the frozen water in polar ice caps and glaciers.**

Water exists in all the *three* physical states : as solid (ice), as liquid (water) and as gas (water vapour).

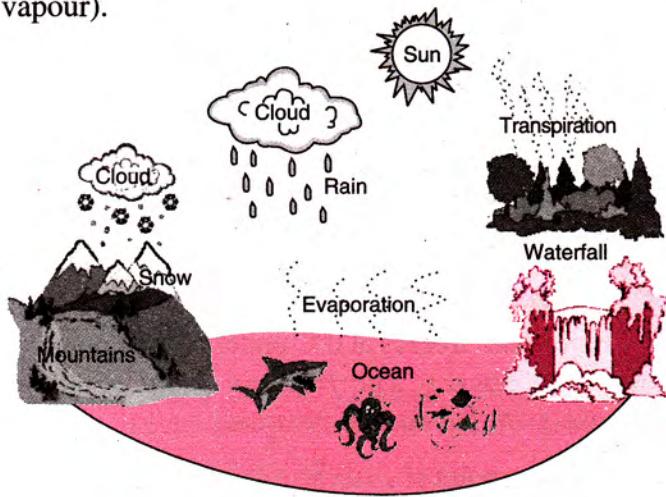


Fig. 3.2 Water cycle.

Water is a major constituent of all living things (plants, animals and human beings) and of the atmospheric environment in which we live. Nearly 70% of our body weight is water.

Water occurs in both free and combined states.

Ordinarily, water exists in the liquid state. When heated under normal pressure (760 mm Hg), water boils at 100°C and changes rapidly into its gaseous state (steam). When steam condenses back into water, heat is released. This heat is called **latent heat of condensation**, and it is equal to latent heat of vaporization.

It is the sudden release of latent heat of condensation that causes the violence associated with *torrential rain*.

### 3.2 PHYSICAL PROPERTIES OF WATER

**1. Nature :** Pure water is a clear, transparent liquid. It is colourless, odourless and tasteless.

The taste in water is due to gases and solids dissolved in it, i.e. impurities present in it.

**2. Boiling point :** Under normal pressure, pure water boils at 100°C. The boiling point of water is affected by pressure because boiling point of a liquid is considered as the temperature when vapour pressure equals the atmospheric pressure. Thus, *the greater the pressure, the higher the boiling point*, and vice versa. This is the principle by which a pressure cooker works.

*In the hills, water boils at temperature lower than 100°C with atmospheric pressure being low (70°C on the top of Mount Everest) and so food is not cooked properly there.* Therefore, pressure cookers are useful in the hills.

The boiling point of water also **increases** due to the presence of dissolved impurities in it.

**3. Freezing point of water or melting point of ice :** Pure water freezes at 0°C under normal pressure (one atmosphere). The *freezing point of water decreases with increase in pressure*.

The freezing point of water also **decreases** due to the presence of dissolved impurities in it.

**4. Density :** At 4°C, water has its maximum density, 1g/cm<sup>3</sup> or 1000 kg/m<sup>3</sup>, and minimum volume. Water expands on freezing, i.e. 92 volumes of water become 100 volumes of ice. Therefore, with relative density of ice being 0.92, it floats.

**5. Anomalous expansion of water :** Water has an unusual physical property. When cooled, it first *contracts* in volume, as do other liquids, but below 4°C, it starts *expanding* and continues to do so till the temperature reaches 0°C, the point at which it freezes into ice.

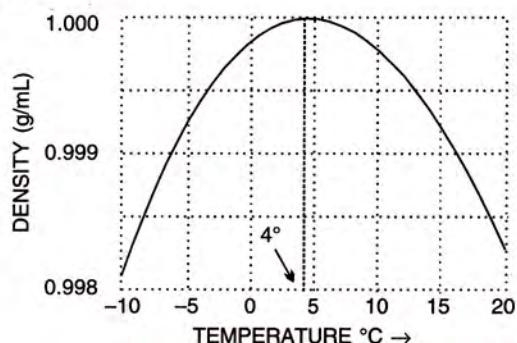


Fig. 3.3 Water has maximum density at 4°C

The property of anomalous expansion of water enables marine life to exist in the colder regions of the world, because even when the water freezes on the surface, it remains a liquid below the ice layer (Fig. 3.4).

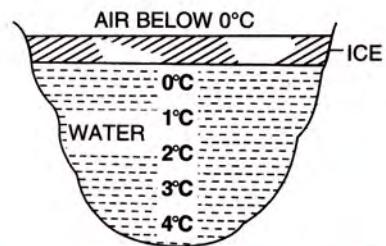


Fig. 3.4 Formation of ice at the top of a pond.

**6. Latent heat of fusion of ice :** The latent heat of fusion is the amount of heat required to change a substance from the solid state to its liquid state at its melting point without any change in temperature.

The amount of heat energy required by ice to change into water, is called **latent heat of fusion of ice**. Its **specific value** is 336 J/g or 80 cal/g. The same amount of heat is released when 1 g of water solidifies to form 1 g of ice at 0°C.

It is on account of high specific latent heat of solidification that lakes and rivers do not freeze suddenly.

**7. Latent heat of vaporization of water :** When water is boiled, it changes to gaseous state.

The energy required to change water into its vapour at its boiling point without any change in temperature is called **latent heat of vaporization of water**, and its **specific value** is 2268 J/g or 540 cal/g. The same amount of heat is released when 1 g of steam condenses to form 1 g of water at 100°C.

It is on account of high specific latent heat of vaporization that steam causes far more serious burns than water at 100°C.

**8. Specific heat capacity :** It has been found that 1 g of water, when heated through 1°C, always absorbs 4.2 J (or 1 calorie) of heat energy. The fixed amount of heat absorbed by 1 g of water, when heated through 1°C, is called its specific heat capacity.

Water because of its high specific heat is commonly used as coolant in motor car radiators, desert coolers, etc.

Due to its high specific heat capacity, the *presence of a large amount of water is able to modify the climate of the nearby land areas*, making them warmer in winter and cooler in summer. Land and sea breeze are also set up because of this great moderating property of water.

### 3.3 WATER AS UNIVERSAL SOLVENT

Water is a remarkable solvent. It dissolves many substances, forming aqueous solutions (water solutions). **Water has a high dielectric constant.** As a result, it reduces the electrostatic force of attraction between positive and negative ions and dissolves even inorganic compounds, which are usually electrovalent. Not only solids, but also gases and other liquids dissolve in water. Hence, *it is difficult to find absolutely pure water in nature. For the same reason, water is also called universal solvent.*

Substances that are apparently *insoluble in water* actually dissolve in it in traces. Even when we put water in a glass vessel, an extremely small amount of glass dissolves in it. It is for this reason that, when *distilled water* is kept in a sealed bottle for a long time, it leaves *etchings on the inside surface of glass*.

**Experiment :** To show that ordinary tap water contains dissolved solids.

**Procedure :** Put some tap water on a clean watch glass, and place it over a beaker containing water, as shown in Fig. 3.5. Boil the water in the beaker. When all the water has evaporated from the watch glass, remove it from the burner and let it cool.

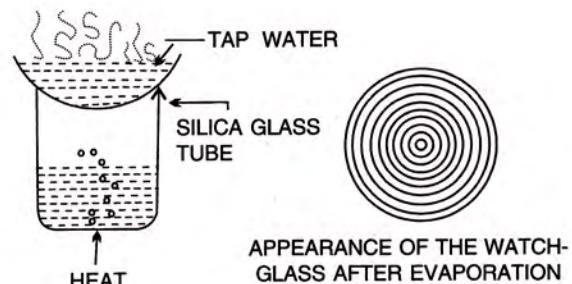


Fig. 3.5 To show that water contains dissolved solids

**Observation :** On looking at the watch glass against light, a number of concentric rings of *solid matter* are seen. These are the dissolved solids left behind after evaporation of water.

### Importance of dissolved salts in water

The dissolved solids in water are salts, minerals and impurities.

- They are essential for growth and development of plants.
- They add taste to water.
- They supply the essential minerals needed by our bodies.

**Note :** Tap water, river water and well water contain dissolved solids, but **rainwater** and **distilled water** do not contain dissolved solids, and so concentric rings are not formed in their case.

### AIR DISSOLVED IN WATER

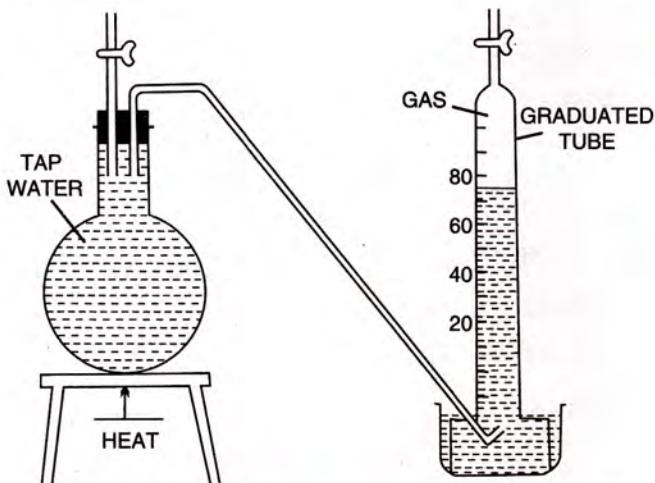
Air is present in dissolved state in all the natural sources of water. Of the two main components of air (nitrogen and oxygen), oxygen is more soluble in water than nitrogen.

**Note :** Since oxygen is more soluble in water compared to nitrogen, the air dissolved in water differs from ordinary air.

**Experiment :** To show that tap water contains dissolved gases (air).

**Procedure :** Fill a round-bottom flask completely with tap water, and arrange the apparatus as shown in Fig. 3.6.

A graduated tube or burette filled completely with water is inverted over the end of the delivery tube in a trough of water.



*Fig. 3.6 To show that water contains dissolved gases*

***Observation :*** Heat the water in the flask. Gas bubbles are seen escaping from water. They are collected in the graduated tube by downward displacement of water. The water vapour present in the gases condenses on coming in contact with the cold water in the tube. Obviously, the remaining gases come from the air dissolved in water, and they escape on heating. Read the volume of the gas thus collected. The volume of the gas (air) collected in the graduated tube shows the presence of dissolved gases in tap water.

The solubility of a gas is often expressed by the volume of that gas dissolved in a certain volume of a solvent. *For example*, the solubility of H<sub>2</sub> in water

is expressed as 0·02 volume of hydrogen per unit volume of water, at 0°C or 506 volumes of hydrogen chloride per unit volume of water, at 0°C.

*The composition of air dissolved in water is 33% oxygen (compared to 21% in ordinary air), 66% nitrogen and 1% carbon dioxide.*

**Note :** The gases dissolved in water can be easily expelled by boiling the water (Fig. 3.6).

**Note :** Distilled water and boiled water have no taste. The pleasant taste of drinking water is due to the presence of dissolved matter, like air, carbon dioxide and minerals.

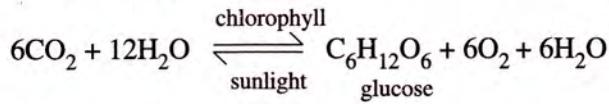
## **IMPORTANCE OF AIR DISSOLVED IN WATER**

Air dissolved in water is biologically very important.

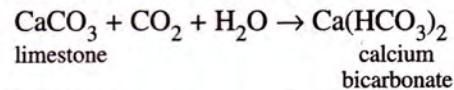
- (1) Marine life, like fish, use the **oxygen** of the air dissolved in water for respiration, and thus aquatic life is sustained. 1 dm<sup>3</sup> of water contains nearly 40 cm<sup>3</sup> of dissolved oxygen.

### (2) Carbon dioxide

- (i) Aquatic plants make use of dissolved **carbon dioxide** for photosynthesis, *i.e.* to prepare their food.



- (ii) Carbon dioxide dissolved in water reacts with limestone to form calcium bicarbonate.



Marine organisms such as snails, oysters, etc. extract calcium carbonate from calcium bicarbonate to build their shells.

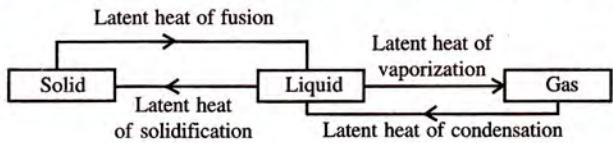
### **EXERCISE 3(A)**

- Water exists in all the three states. Discuss.
  - Why is water considered a compound ?
  - (a) Why does temperature in Mumbai and Chennai not fall as low as it does in Delhi ?  
(b) Give the properties of water responsible for controlling the temperature of our body.
  - ‘Water is a universal solvent’. Comment.
  - What causes the violence associated with torrential rain ?
  - (a) Which property of water enables it to modify the climate ?  
(b) Density of water varies with temperature. What are its consequences ?  
(c) State the melting point, boiling point, specific heat capacity, specific latent heat of fusion and specific latent heat of vaporization of water.
  - How does fishes and aquatic animals survive when the pond gets covered with thick ice ?

8. The properties of water are different from the properties of the elements of which it is formed. Discuss.
9. How is aquatic life benefitted by the fact that water has maximum density at  $4^{\circ}\text{C}$  ?
10. Describe an experiment to show :
  - (a) tap water contains dissolved salts.
  - (b) tap water contains dissolved gases.
11. State the importance of the solubility of  $\text{CO}_2$  and  $\text{O}_2$  in water.
12. How is air dissolved in water different from ordinary air ?
13. Why do rivers and lakes not freeze easily ?
14. What is the importance of dissolved salts in water ?
15. Explain why :
  - (a) Boiled or distilled water taste flat.

- (b) Ice at zero degree centigrade has greater cooling effect than water at  $0^{\circ}\text{C}$ .
- (c) Burns caused by steam are more severe than burns caused by boiling water.
- (d) Rain water does not leave behind concentric rings when boiled.
- (e) Air dissolved in water contains a higher proportion of oxygen.
- (f) If distilled water is kept in a sealed bottle for a long time, it leaves etchings on the surface of the glass.

16. Explain what you understand from the following diagram :



## 3 B. SOLUTIONS

### 3.4 SOLUTIONS AS MIXTURES OF SOLIDS IN WATER

A solution is a **homogeneous mixture** of two or more components whose composition may be gradually changed by changing the relative amounts of the components.

$$\text{Solution} = \text{Solute} + \text{Solvent}$$

The medium of dissolution that allows one or more component to dissolve in it to form a solution is known as **SOLVENT**. The substance that dissolves in the solvent to form a solution is known as **SOLUTE**.

A solution which is made up of two components is a **binary** solution. Solutions which are made up of three or four components are called **ternary** and **quaternary** solutions respectively. Here we shall deal with binary solutions only, composed of a solute and a solvent.

*For example :* When salt crystals dissolve in water to form saline water, the salt is classified as the **solute**, the water is classified as the **solvent** and the saline water is an example of a **solution**.

If the components exist in different physical states, then the solvent is that component which exists in the same phase (state) as the final solution.

*For example,* if sugar is dissolved in water, the final solution is a liquid solution. So water (liquid) is the solvent and sugar the solute.

If the two substances are in the same phase (state), the component that is in excess is termed the **solvent**.

A solution may be formed by mixing :

- (i) a liquid with another liquid

*Example :* Water and alcohol

- (ii) a solid with a liquid

*Example :* Sugar solution

- (iii) a liquid with a gas.

*Example :* Soda water

- (iv) a solid with another solid

*Example :* Alloys

A **homogeneous** solution of a solid into another solid is called a **solid solution**. Common metal alloys are solid solutions. The gold used in making jewellery is a solid solution containing gold and small quantities of copper or silver.

Brass is a solid solution containing 70% copper and 30% zinc. Here, copper acts as the solvent and zinc as the solute.

A solution in which the size of the solute particles is about  $10^{-10}$  m is called a **true solution**. In a true solution, the solute particles and the solvent molecules cannot be distinguished even under a microscope. In common use, **the term solution actually means true solution**.

## Characteristics of a true solution (or simply a solution)

1. It is a homogeneous mixture.
2. The solute particles are very small, about  $10^{-10}$  m.
3. It is clear and transparent.
4. It does not scatter light.
5. Its components cannot be separated by filtration.
6. The solute particles in a solution do not settle down.

### Note :

1. Water is the most common solvent (a solution produced by dissolving a substance in water is known as an **aqueous solution**) but it is not the only solvent. Alcohol, petrol, ether, benzene, carbon disulphide, liquid ammonia, etc. are some **non-aqueous** liquid solvents in common use. The solutions made in these liquids are known as non-aqueous solutions.
2. When the solvent is alcohol, the solution is called a **tincture**. *Tincture of iodine is a solution of iodine in alcohol.*

### Dilute solution

It is the solution in which the amount of solute is rather small compared to the mass of the solvent.

### Concentrated solution

It is the solution in which the amount of solute is relatively large for a given mass of the solvent.

## 3.5 SATURATED SOLUTION

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

### Preparation of a saturated solution

**Experiment :** Take 100 g of distilled water in a beaker. Add 1 g of nitre and stir with a glass rod till the nitre dissolves. Add one gram more of nitre and stir it. It too will dissolve. Continue adding a gram of nitre at a time, and stir vigorously after each addition. A stage will come when no more of the nitre will dissolve at that temperature, no matter how long it is left there or how vigorously it is stirred. The solution is then said to be **saturated** with nitre at that temperature. This is the saturated solution of

nitre in water and it cannot dissolve any more of the solute without changing temperature.

A **solution** is a homogeneous mixture of a solute in a solvent; its composition may vary till a saturated solution is obtained at a particular temperature.

### Effect of cooling on a saturated solution

The solubility of a solid usually decreases with fall in temperature. Thus, if the temperature of a saturated solution is lowered, or if a solution is left open at a higher temperature for long, (the solvent slowly vaporizes) and a part of the dissolved solute separates out in the form of crystals. The remaining solution is still a saturated solution but at a lower temperature.

A saturated solution can also be made unsaturated by adding more solvent to it.

### Unsaturated solution

A solution in which more of the solute can be dissolved at a given temperature is known as an **unsaturated solution**.

### Supersaturated solution

A solution that holds more solute than it can theoretically do at room temperature is a supersaturated solution. Heat a saturated solution and dissolve some more solute to it. Then cool it back to room temperature; you will get a supersaturated solution since it holds more solute than it can theoretically hold at that temperature.

A supersaturated solution at a particular temperature is one that is more concentrated (contains more solute) than its saturated solution at that temperature.

### Preparation of supersaturated solutions

A saturated solution of a solute, say potassium nitrate, is prepared in slightly warm water. On cooling this solution, the excess nitre does not separate out. This solution contains more solute in it than it can hold at room temperature, and therefore it is a **supersaturated solution**.

**Note :** It is possible to bring down the temperature of a saturated solution a little without the excess solute crystallizing out, if :

- no suspended impurities, like dust particles, are present in it.
- the solution is not disturbed physically, i.e. it is left standing still in its container.

### 3.6 CONCENTRATION OF A SOLUTION

Concentration of a solution is the amount of solute dissolved in a given quantity of that solution. It can be expressed in many ways. *For example :*

#### (i) Mass percent

It is defined as the mass of solid solute in grams present in 100 grams of the solution. It is mostly used when solute is solid and solvent is liquid.

#### Concentration of solution

$$= \frac{\text{Mass of solute}}{\text{Mass of solution (solute + solvent)}} \times 100$$

*For example,* if 10 grams of sodium chloride is added to 90 grams of water, the concentration  
 $= \frac{10}{10+90} \times 100 = \frac{10}{100} \times 100 = 10\%$ .

#### (ii) Volume percent

It is defined as the volume of solute in millilitres present in 100 mL of a solution. It is mostly used when solute and solvent are liquid.

#### Volume percent

$$= \frac{\text{Volume of solute}}{\text{Volume of solute} + \text{Volume of solvent}} \times 100$$

*For example,* if 30 mL of alcohol is mixed with 70 mL of water to get 100 mL of solution, volume percent

$$= \frac{30}{30+70} \times 100 = \frac{30}{100} \times 100 = 30\%$$

#### Solved examples :

*Example 1 :* 2.5 litres of alcohol is present in 10 litres of aqueous solution of alcohol. Calculate volume percent.

$$\text{Volume of solute} = 2.5 \text{ litres}$$

$$\text{Volume of solution} = 10.0 \text{ litres}$$

**Solution :** Volume percent

$$= \frac{\text{Volume of Solute}}{\text{Volume of solution}} \times 100 \\ = \frac{2.5}{10} \times 100 = 25\%$$

**Example 2 :** 50 gram of sugar is dissolved in 2.45 kg of water. Calculate the concentration of solution.

**Solution :** Mass of solute = 50 gram,

mass of solvent = 2450 gram,

mass of solution = mass of solute + mass of solvent = 2500 gram

$$\text{Mass\%} = \frac{\text{Mass of solute} \times 100}{\text{Mass of solution}}$$

$$\text{Mass\%} = \frac{50}{2500} \times 100 = 2\% \text{ solution}$$

### 3.7 SOLUBILITY

Different substances dissolve to different extents in the same volume of a given solvent. In other words, they have different solubilities. *For example,* the masses of sodium chloride and potassium nitrate that dissolve in 100 g of water at 40°C are 36.5 g and 65 g respectively. Obviously, the solubility of potassium nitrate is greater than that of sodium chloride.

The solubility of a solute in a particular solvent at a given temperature is equivalent to the maximum number of grams of the solute necessary to saturate 100 g of that solvent at that temperature.

A substance that has negligible solubility is called **insoluble**, e.g., silver chloride, which has 0.000015 g as its solubility. On the other hand, a substance that has more than negligible but less than high solubility is called **sparingly soluble**, e.g. calcium hydroxide, which has 0.17 g as its solubility. A substance having high solubility such as sodium chloride is simply called **soluble**.

#### Determination of the solubility of a solute at a particular temperature

- Note the temperature of the saturated solution.
- Weigh a clean, dry evaporating dish = M.
- Put the saturated solution in it and weigh again = M<sub>1</sub>.

- Now heat the solution to dryness and again weigh =  $M_2$ .
- Mass of saturated solution =  $M_1 - M$ .
- Mass of solute =  $M_2 - M$ .
- Mass of solvent = Mass of saturated solution – mass of solute =  $(M_1 - M) - (M_2 - M)$

At the noted temperature :

$$\begin{aligned}\text{Solubility} &= \frac{\text{Mass of solute}}{\text{Mass of solvent}} \times 100 \\ &= \frac{M_2 - M}{(M_1 - M) - (M_2 - M)} \times 100\end{aligned}$$

### 3.7.1 Factors affecting solubility

The rate of dissolution or rate of solubility of a solid in a liquid depends on the following factors :

- (a) **Size of solute particles** : The smaller the size of the solute particles, the greater is its total surface area exposed to the solvent. Therefore, the greater is the solubility of that solute.
- (b) **Stirring** : This brings more of the solvent in contact with the solute and thus increases the rate of formation of solution.
- (c) **Temperature** : The solubility of a gas in a liquid always decreases with rise in temperature. But **the solubility of most solids in water usually increases with rise in temperature**.

For example, the solubility of potassium nitrate in water at  $20^\circ\text{C}$  is 31.6 grams, whereas its solubility at  $60^\circ\text{C}$  is 63.9 grams.

All these factors affect solubility because in process of dissolution, the particles of solute merely occupies the spaces between particles of solvent without undergoing chemical change.

### 3.7.2 Solubility curves

A *solubility curve* is a line graph that plots changes in solubility of a solute in a solvent against changing temperature.

Variation in solubility of a given substance with temperature can be shown by means of *solubility curves*. Here, solubility is plotted along the ordinate (y-axis) and temperature along the abscissa (x-axis). Solubility curves of some common compounds are shown in Fig. 3.7.

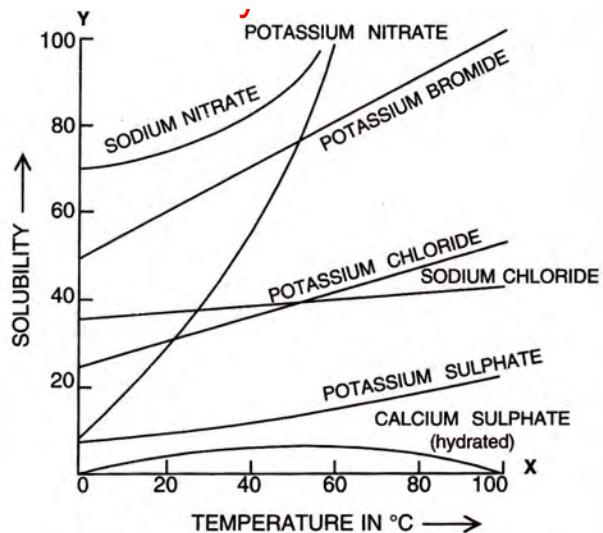


Fig. 3.7 Variation in solubility of some solids with changing temperature

- Solubility curves of substances like **calcium sulphate** ( $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ ) show that there is a decrease (after attaining a certain temperature) in their solubilities with further rise in temperature. Other examples are ceric sulphate  $\text{Ce}_2(\text{SO}_4)_3$ , slaked lime  $\text{Ca}(\text{OH})_2$ , etc.
- Substances like **sodium nitrate**, **potassium nitrate**, **potassium bromide**, etc., show considerable increase in their solubilities with rise in temperature.
- Solubility of **sodium chloride** increases only a little with increase in temperature.

In an **endothermic process** the solubility of solute increase with the increase of temperature. For example solubility  $\text{KNO}_3$  increases with rise in temperature. In an enothermic process, the solubility increases on lowering the temperature. For example, solubility by calcium hydroxide in water decreases on increasing the temperature.

#### Anomalous solubility :

There are some salts whose solubilities first increase, and then decrease, with rise in temperature.

For example :  $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$  (Glauber salt)

Solubility curve of  $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$  rises till it reaches  $32.8^\circ\text{C}$ , and then it falls slightly. This is because  $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$  is hydrous below  $32.8^\circ\text{C}$  and anhydrous above it (Fig. 3.8).

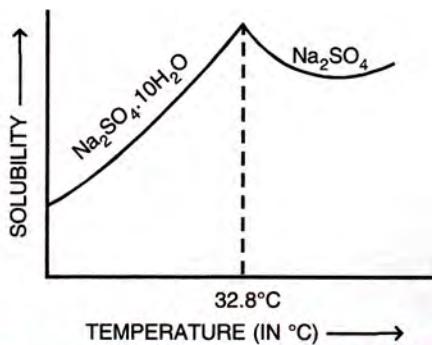


Fig. 3.8 Solubility curve for  $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$

### 3.7.3 Applications of solubility curves

- Shape of the curve indicates how the solubility of the given substance in a solvent varies with change in temperature. The solubility of a substance at a particular temperature can be determined from the curve.
- Solubilities of different solutes in a solvent at different temperatures can be compared by comparing at their respective solubility curves plotted on the same chart (Fig. 3.7).
- The effect of cooling of hot solutions of different substances can be found from the curves.

When a saturated solution is prepared at  $40^\circ\text{C}$  and then it is cooled to  $15^\circ\text{C}$ , some salts gets precipitated out. On the other hand, if a saturated solution is heated to a higher temperature, then it becomes unsaturated. More solute can be dissolved in this solution now.

### Some solved examples :

#### Example 1 :

12 g of a saturated solution of potassium chloride at  $20^\circ\text{C}$ , when evaporated to dryness, leaves a solid residue of 3 g. Calculate the solubility of potassium chloride.

#### Solution :

$$\text{Weight of water in solution} = 12 \text{ g} - 3 \text{ g} = 9 \text{ g}$$

9 g of  $\text{H}_2\text{O}$  dissolves 3 g of solid

$\therefore$  100 g of water will dissolve

$$\frac{3}{9} \times 100 = 33.3 \text{ g}$$

**Ans.** Solubility of KCl in  $\text{H}_2\text{O}$  at  $20^\circ\text{C}$  = 33.3 g

#### Example 2 :

Find the weight of sodium nitrate required to prepare 60 g pure crystals from its saturated solution at  $70^\circ\text{C}$ . Solubility of sodium nitrate is 140 g at  $70^\circ\text{C}$  and 100 g at  $25^\circ\text{C}$ .

#### Solution :

Solubility at  $70^\circ\text{C}$  = 140

Solubility at  $25^\circ\text{C}$  = 100

Amount of crystals obtained when the solution is cooled from  $70^\circ\text{C}$  to  $25^\circ\text{C}$

$$= 140 - 100 = 40 \text{ g}$$

To obtain 40 g of crystals, sodium nitrate taken is 140 g.

To obtain 60 g crystals, sodium nitrate required will be  $\frac{140}{40} \times 60 = 210 \text{ g}$

**Ans.** 210 g

### 3.7.4 Effect of pressure and temperature on solubility of gases in water (liquids)

**Pressure :** An increase in pressure on the surface of water increases the solubility of a gas in water.

At any given temperature, the mass of gas dissolved by a fixed volume of liquid is directly proportional to the pressure on the surface of the liquid. This law is known as **Henry's Law**.

For example, the solubility of carbon dioxide in water under normal atmospheric pressure is rather low, but when the water surface is subjected to higher pressure a lot more of  $\text{CO}_2$  gas dissolves in it, as is seen in the case of **soda water**. On opening the soda water bottle, the dissolved gas rapidly bubbles out since pressure on the surface of water suddenly decreases.

**Temperature :** An increase in temperature of water causes a decrease in solubility of a gas in it. Thus, on boiling, water loses its taste. Since some taste of water is due to the gases in it, on boiling, these gases escape from water, leaving it tasteless.

### 3.8 CRYSTALS AND CRYSTALLIZATION

A **crystal** is a homogeneous solid of definite geometrical shape. It has symmetrically arranged smooth plane surfaces which meet forming the sharp edges.

**Crystallization** is a process by which crystals of a substance are obtained by cooling a hot saturated solution.

All crystals of a pure compound are of similar shape, but those of different compounds may be of different shapes. Some shapes in which crystals exist are :

Cubic (e.g.,  $\text{NaCl}$ )

Rhombohedral (e.g.,  $\text{CuSO}_4$ )

Octahedral (e.g.,  $\text{FeSO}_4$ )

Prismatic, with sharp edges (e.g.,  $\text{KNO}_3$ )

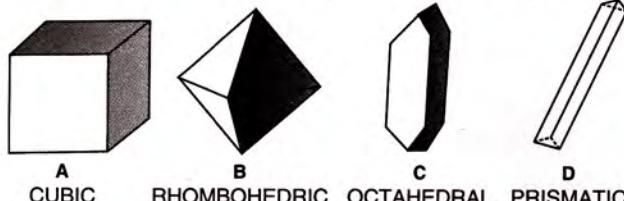


Fig. 3.9 Some shapes of crystals

In nature too, several crystals are formed, some of which are recovered from the earth as minerals, e.g. common salt, potassium nitrate, ruby, saphire, diamond. In some cases, their beautiful colours are due to traces of water or impurities present in them. For example, the red colour of ruby is due to a trace of mercury oxide present in its crystals.

In a laboratory, crystals may be obtained by the following methods :

- (1) By cooling a hot saturated solution
- (2) By slowly evaporating a saturated solution
- (3) By cooling a fused mass
- (4) By sublimation

**Experiment :** To prepare large crystals of copper sulphate.

**Procedure :** Make a saturated solution of copper sulphate at  $80^\circ\text{C}$ . Filter it. Allow the filtrate to cool and leave it for a day. Next morning, a number of tiny crystals will be seen lying at the bottom of the beaker. Pick up these small crystals, which are well formed. Now, suspend a well-shaped crystal by means of a thread tied to a glass rod into the cooled saturated solution (Fig. 3.10). Cover the beaker with a piece of paper to prevent dust from getting in. Leave the beaker undisturbed and watch the crystal grow.

Table 3.2 Some salts containing water of crystallisation

Common name	Chemical name	Formula
Washing soda crystals	Sodium carbonate decahydrate	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
Epsom salt	Magnesium sulphate heptahydrate	$\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$
Potash alum	Hydrated potassium aluminium sulphate	$\text{K}_2\text{SO}_4 \cdot \text{Al}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$
Glauber's salt	Sodium sulphate decahydrate	$\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$
Blue vitriol	Copper (II) sulphate pentahydrate	$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
Plaster of Paris	Calcium sulphate semihydrate	$\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$
White vitriol	Zinc sulphate heptahydrate	$\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$
Gypsum	Hydrated calcium sulphate	$\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$

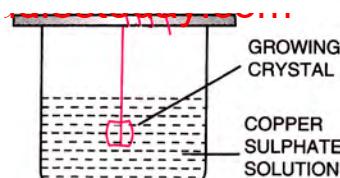


Fig. 3.10 A growing crystal of  $\text{CuSO}_4$

**Observation :** After a few days, the crystal will be found to have grown quite big. The suspended crystal acts as a *seed crystal*.

This process of inducing crystallization by adding a crystal of a pure substance into its saturated solution is called **seeding**.

**Caution :** Avoid addition of dust particles, or crystals will form around them as well.

### 3.9 HYDRATED AND ANHYDROUS SUBSTANCES

Substances which contain water molecules along with salts like sodium carbonate decahydrate ( $\text{NaCO}_3 \cdot 10\text{H}_2\text{O}$ ) and copper sulphate pentahydrate ( $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ ) are hydrated substances.

Substances which do not contain any water along with the salts are anhydrous substances. For example sodium chloride ( $\text{NaCl}$ ), potassium nitrate ( $\text{KNO}_3$ ) etc.

#### 3.9.1 Hydrated substances

The crystals of some salts contain water of crystallisation. This water gives the crystals their shape. In some cases it also gives them their colour (copper sulphate crystals are blue in colour). Such salts are known as **hydrated salts**.

#### WATER OF CRYSTALLIZATION

The fixed amount of water that is associated with hydrated crystals which is an integral part of the crystal, is called **water of crystallisation**.

Some solids, while crystallizing out from their solutions, combine with a definite quantity of water, known as water of crystallization or water of hydration.

This water of crystallization is in loose chemical combination and can be driven off as steam by heating the crystals above 100°C, the vapours produced are **condensed and the liquid formed is tested**. This liquid gives test of water. In this process, the crystals lose their water and become **anhydrous**, i.e. all the water combined in the crystalline material is removed.

**Experiment :** To show that hydrated copper sulphate crystals (blue vitriol) contain water of crystallization.

**Procedure :** Take some powdered copper sulphate crystals in a clean dry test tube, and clamp it in a tilted position (Fig. 6.11) so that the condensed water does not trickle or flow back. Gently heat the crystals.

**Observation :** Drops of colourless liquid condense on the cooler parts of the test tube, leaving behind a residue that is anhydrous (without water) and amorphous (non-crystalline), i.e. with no definite shape or structure.

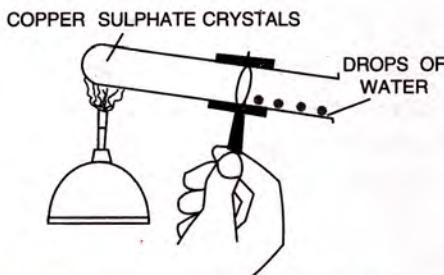
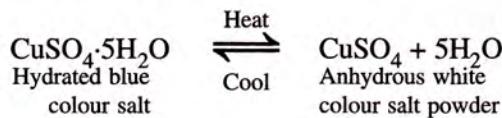


Fig. 3.11 To show loss of water of crystallization on heating

When blue crystals of hydrated copper sulphate are heated in a test tube, they turn into a white powder, which turns back into a blue solid when a few drops of water are added.



When preparing crystals of a hydrated salt, we must be careful not to heat them too strongly when drying them, otherwise the product is a dehydrated powder, not crystals.

### Determination of water of crystallization

Take a known weight of crystals in a china dish

and heat it above 100°C. When the weight of the residue becomes constant, stop heating.

Wt. of crystals at room temperature =  $a$  g

Wt. of crystals after heating at 101°C =  $b$  g

Wt. of water =  $(a - b)$  g

Now  $a$  g of crystals contain water =  $(a - b)$  g

$$\therefore \% \text{ of water of crystallization} = \frac{a - b}{a} \times 100$$

### 3.9.2 Anhydrous substances

A substance is anhydrous if it contains no water, for example, salts lacking their water of crystallisation. The term 'Anhydrous' is most often applied to crystalline substances when the water of crystallisation is removed.

Anhydrous can also refer to the gaseous form of some concentrated solutions such as ammonia to distinguish it from its aqueous solution form.

Anhydrous solvents are used to perform certain chemical reactions that cannot proceed in presence of water. Examples of anhydrous substances.

- Table salt is anhydrous sodium chloride.
- Gaseous HCl is anhydrous.

Anhydrous salts can be obtained from hydrated salts by carefully removing the water of crystallisation from them. It can be done by

- Direct heating.
- Heating in dry and hot air.
- Heating under vacuum.
- By using dehydrating/desiccating agents such as warm concentrated sulphuric acid.

The crystalline shape of a substance is not necessarily the result of the presence of water of crystallization. In fact, there are a number of crystalline solids that crystallize from water without holding any water of crystallization.

Examples of crystalline substance which do not contain water of crystallization :

Common salt (NaCl),

Nitre (KNO<sub>3</sub>),

Sugar (C<sub>12</sub>H<sub>22</sub>O<sub>11</sub>),

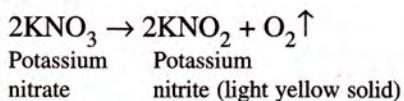
Potassium permanganate (KMnO<sub>4</sub>),

Ammonium chloride (NH<sub>4</sub>Cl).

**Experiment :** To study the action of heat on anhydrous crystals of sodium chloride and potassium nitrate.

**Procedure :** Gently heat a few crystals of pure sodium chloride (common salt) in one test tube and a few crystals of potassium nitrate in another.

**Observation :** On heating, crystals of sodium chloride produce a crackling sound, called **decrepitation**. This is due to the breaking of bigger crystals of sodium chloride into smaller ones. On the other hand, crystals of potassium nitrate first melt into a colourless liquid, and then, on strong heating, they produce oxygen, which rekindles a glowing splint. Potassium nitrite, a light yellow solid residue, is left behind.



**Conclusion :** Both the substances do not contain water of crystallization.

## 3.10 PROPERTIES

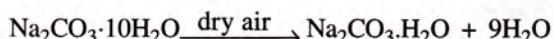
### 3.10.1 Efflorescence

A phenomenon where a compound loses its water of crystallisation on exposure to dry air is called **efflorescence**. This results in loss of crystalline shape, and finally the crumbling of crystals into a powder. Substances showing **efflorescence** are called **efflorescent substances**.

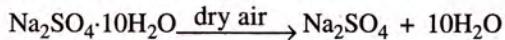
Efflorescence occurs when vapour pressure in the hydrated crystals exceeds atmospheric vapour pressure. Thus, it is minimized during humid conditions.

#### Examples of efflorescent substances :

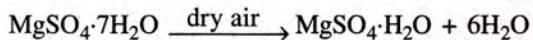
(a) **Washing soda** [hydrated sodium carbonate], when exposed to dry air, becomes a monohydrate.



(b) **Glauber's salt** [hydrated sodium sulphate,  $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ ] becomes a powdery anhydrous sodium sulphate when exposed to air.



(c) **Epsom salt** [magnesium sulphate heptahydrate], when exposed to dry air, becomes a monohydrate.



**Note :** The higher the temperature of the air, the higher the efflorescence. This is because the air absorbs more water with rising temperature and decreasing moisture.

### 3.10.2 Deliquescence

Certain water soluble substances, when exposed to the atmosphere at ordinary temperatures, absorb moisture from atmospheric air, become moist, lose their crystalline form, and ultimately dissolve in the absorbed water, forming a saturated solution. Such a substance is called a **deliquescent substance** and the phenomenon is called **deliquescence**.

Deliquescence occurs when vapour pressure inside the crystals is very low compared to vapour pressure in the **atmospheric air**. Thus, deliquescence is minimized during dry conditions.

#### Deliquescent substances

Example of caustic soda  $\text{NaOH}$ , caustic potash  $\text{KOH}$ , magnesium chloride  $\text{MgCl}_2$ , zinc chloride  $\text{ZnCl}_2$ , calcium chloride  $\text{CaCl}_2$  and ferric chloride  $\text{FeCl}_3$ .

**Note :** Table salt [sodium chloride] turns moist and ultimately forms a solution, on exposure to air [especially during the rainy season]. Though pure sodium chloride is not deliquescent, the commercial version of the salt contains **impurities**, like **magnesium chloride** and **calcium chloride**, which are deliquescent substances.

This kind of impurity can be removed by passing a current of dry hydrogen chloride gas through a saturated solution of the affected salt. Pure sodium chloride is produced as a precipitate, which can be recovered by filtering and washing first with a little water and finally with alcohol.

### 3.10.3 Hygroscopic Substances

Certain substances absorb moisture (water vapour) from the atmosphere when they are exposed to it. Such substances are called **hygroscopic substances**. The phenomenon is known as **hygroscopy**.

Some examples of hygroscopic substances are :

- Conc. sulphuric acid ( $\text{H}_2\text{SO}_4$ )
- Phosphorus pentoxide ( $\text{P}_2\text{O}_5$ )
- Quicklime ( $\text{CaO}$ )

Hygroscopic substances have a tendency to take up moisture from the air.

### 3.11 DRYING AND DEHYDRATING AGENTS

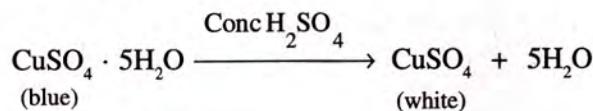
Drying agents are substances that can readily absorb moisture from other substances without chemically reacting with them.

For example, anhydrous calcium chloride, anhydrous zinc chloride, phosphorus pentoxide, magnesium sulphate, dry sodium sulphate, etc., are used to absorb water vapour from the air.

These substances are also called **desiccants** or **desiccating agents**.

Almost all hygroscopic substances are desiccating agents, e.g. conc. sulphuric acid, phosphorus pentoxide, silica gel, quicklime.

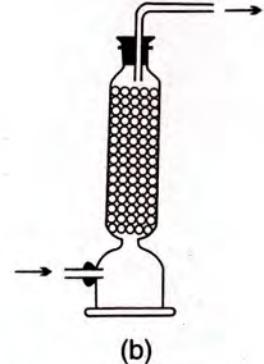
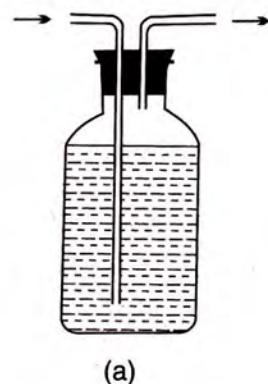
Dehydrating agents are substances that can remove even the chemically combined water molecules from compounds. For example, conc. sulphuric acid can remove water molecules from blue vitriol ( $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ ). Therefore, it is a dehydrating agent.



#### Gases are dried by :

(i) Passing the gases through conc. sulphuric acid  
conc. sulphuric acid has strong affinity for water, so it easily extracts water from many substances [Fig. 3.12(a)].

Conc. Sulphuric acid is used to dry gases which



are acidic in nature like hydrogen chloride gas.

- Passing them through a drying tower or a U-tube containing anhydrous sodium sulphate [Fig. 3.12(b) and Fig. 3.12(c)].
- A drying bulb containing anhydrous calcium chloride [Fig. 3.12(d)].

Quicklime being basic in nature is suitable for drying  $\text{NH}_3$ , a basic gas. Basic gas will not react with a base.

Liquids are dried by keeping them overnight over anhydrous  $\text{Na}_2\text{SO}_4$  or  $\text{MgSO}_4$  or  $\text{CaCl}_2$  at room temperature. The solid is then removed by filtration.

Solids are dried by spreading them on a watch glass or a dish and keeping it in a desiccator for some days.

A desiccator is an air-tight glass vessel with a suitable drying agent (*anhydrous calcium chloride*) placed at the bottom. It is used for drying the solids placed in it.

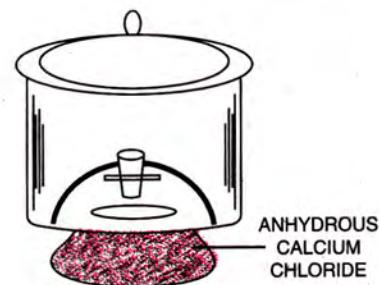


Fig. 3.13 A desiccator

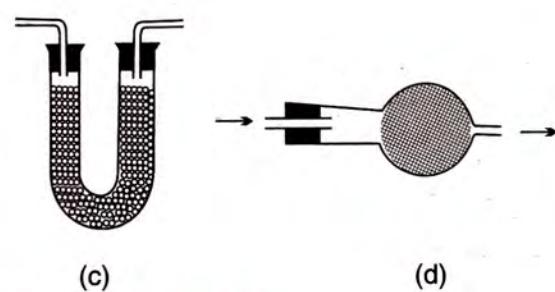


Fig. 3.12 (a) Drying bottle containing concentrated sulphuric acid (b) drying tower (c) U-tube (d) drying bulb

**Drying agent**

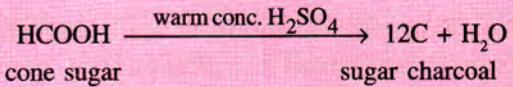
- (i) They remove moisture from other substances.
- (ii) They are used to dry gases like chlorine, sulphur dioxide, hydrogen chloride, etc. They are also used in desiccators to keep substances dry.
- (iii) They represent physical change.

**Examples :**

Phosphorus pentoxide  $P_2O_5$ , fused calcium chloride  $CaCl_2$ , calcium oxide  $CaO$ , conc. sulphuric acid  $H_2SO_4$ .

**Dehydrating agent**

- (i) They remove chemically combined elements of water in the ratio of 2 : 1 (hydrogen : oxygen) from a compound.
- (ii) They prepare substances like carbon monoxide, sugar charcoal, etc.



- (iii) They represent chemical change.

**For example :**

Conc. Sulphuric acid  $H_2SO_4$ .

**EXERCISE 3(B)**

1. Explain the terms :
    - (a) solution (b) solute (c) solvent.
  2. Explain why a hot saturated solution of potassium nitrate forms crystals as it cools.
  3. Give *three* factors which affect the solubility of a solid solute in a solvent.
  4. If you are given some copper sulphate crystals, how would you proceed to prepare its saturated solution at room temperature ? Give practical details. How can you show that your solution is really saturated ?
  5. (a) Define (i) Henry's law (ii) Crystallization.  
 (b) State the different methods of crystallization.
  6. What would you observe when crystals of copper (II) sulphate and iron (II) sulphate are separately heated in two test tubes ?
  7. Table salt becomes sticky on exposure to humid air during the rainy season. Explain.
  8. What is the effect of temperature on solubility of  $KNO_3$  and  $CaSO_4$  in water ?
  9. Solubility of  $NaCl$  at  $40^\circ C$  is 36.5 g. What is meant by this statement ?
  10. Which test will you carry out to find out if a given solution is saturated or unsaturated or supersaturated ?
  11. What is the effect of pressure on solubility of gases. Explain with an example.
  12. What are solubility curves ? Give their uses.
  13. Explain why :
    - (a) water is an excellent liquid to use in cooling systems.
    - (b) a solution is always clear and transparent.
    - (c) lakes and rivers do not suddenly freeze in the winters.
    - (d) the solute cannot be separated from a solution by filtration.
  - (e) fused  $CaCl_2$  or conc.  $H_2SO_4$  is used in a desiccator.
  - (f) effervescence is seen an opening a bottle of soda water.
  14. Normally, solubility of a crystalline solid increases with temperature. Does it increase uniformly in all cases ? Name a substance whose solubility :
    - (a) increases rapidly with temperature.
    - (b) increases gradually with temperature.
    - (c) increases slightly with temperature.
    - (d) initially increase than decreases with rise in temperature.
  15. What are drying or desiccating agents. Give examples.
  16. Complete the following table :
- | <i>Common Name</i>    | <i>Chemical Name</i> | <i>Formula</i> | <i>Acid, base or salt</i> | <i>Efflorescent, hygroscopic or deliquescent substance</i> |
|-----------------------|----------------------|----------------|---------------------------|--|
| Solid caustic pottash |                      |                |                           |  |
| Quick lime            |                      |                |                           |  |
| Oil of vitriol        |                      |                |                           |  |
| Washing soda          |                      |                |                           |  |
| Solid caustic soda    |                      |                |                           |  |
| Blue vitriol          |                      |                |                           |  |

17. In which of the following substances will there be :  
 (a) increase in mass      (b) decrease in mass  
 (c) no change in mass when they are exposed to air ?
1. Sodium chloride
  2. Iron
  3. Conc. sulphuric acid
  4. Table salt
  5. Sodium carbonate crystals
18. To make a saturated solution, 136 g of a salt is dissolved in 500 g of water at 293 K. Find its solubility at this temperature. **Ans.** 27.2g
19. (a) A solution contains 15 g of sodium chloride in 285 g of water. Calculate concentration of solution. **Ans.** 5%
- (b) 4 litres of an organic compound, acetone, is present in 90 litres of an aqueous solution. Calculate its volume percent. **Ans.** 4.44%
20. The following table gives the solubility at different temperatures of different salts,

	Temperature in Kelvin				
Substance dissolved	283	293	313	333	353
KNO <sub>3</sub>	21	32	62	106	167
NaCl	36	36	36	37	37
KCl	35	35	40	46	54
NH <sub>4</sub> Cl	24	37	41	55	66

Answer the following questions based on the table given above.

- (a) What mass of KNO<sub>3</sub> would be needed to produce a saturated solution of KNO<sub>3</sub> in 50 grams of water at 313 K.
- (b) If saturated solution of KCl is made at 353 K and then cooled at room temperature, what would you observe? Explain.
- (c) Find the solubility of each salt at 293 K.
- (d) Which salt has the lowest solubility at 283 K ?
- (e) What is the effect of change of temperature on the solubility of a salt.
- Ans.** (a) 31 g (b) Crystals of KCl (c) KNO<sub>3</sub> 32 g, NaCl 36 g, KCl 35 g, NH<sub>4</sub>Cl 37 g (d) KNO<sub>3</sub>
21. (a) Find the solubility of KNO<sub>3</sub> at 20°C when the mass of the empty dish is 50 g, the mass of dish and solution is 65 g, while the mass of dish and residue is 54.3 g. **Ans.** 40.18 g
- (b) What weight of sodium nitrate will separate when a saturated solution containing 50 g of water is cooled from 50°C to 30°C ? The solubility of NaNO<sub>3</sub> at 50°C and 30°C is 114 g and 86 g respectively. **Ans.** 14 g
22. To make a saturated solution, 36 g of sodium chloride is dissolved in 100 g of water at 293 K. Find its concentration at this temperature. **Ans.** 26.47%

### 3 C. SOFT AND HARD WATER

#### 3.12 SOFT AND HARD WATER

Water is said to be soft, if it readily **forms lather with soap**. Pure water or water containing sodium salts easily gives lather with soap. Such a water sample is called 'soft water' and is suitable for use. Distilled water and rain water are examples of soft water.

Water is said to be hard when it does not readily form lather with soap.

##### 3.12.1 Causes of hardness

Water of some springs, wells and rivers contain dissolved mineral matter. If this happens to be the

hydrogen carbonates, sulphates or chlorides of calcium and magnesium, the water gets hard.

*Hard water contains bicarbonates, chlorides or sulphates of calcium or magnesium.*

When water flows over beds of limestone and dolomite (CaCO<sub>3</sub>, MgCO<sub>3</sub>) carbon dioxide slowly converts some of these into soluble calcium and magnesium bicarbonates which get mixed with the water and make it hard. The formation of stalactites and stalagmites in caves is due to the same reason.

**Stalagmites and stalactites :** In some lime-stone caves, one sometimes sees conical pillar like objects, some hanging from the roof of the caves and some rising from their floors. These are formed by water dropping from the cracks in the rocks, containing dissolved calcium hydrogen carbonate. Release of pressure results in the conversion of some hydrogen carbonate to calcium carbonate.



This calcium carbonate, little by little gets deposited both on the roof and floor of the cave.

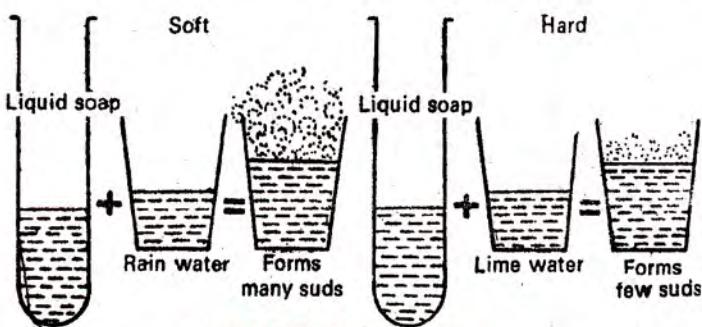
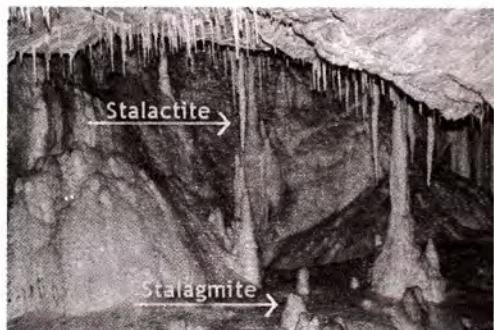


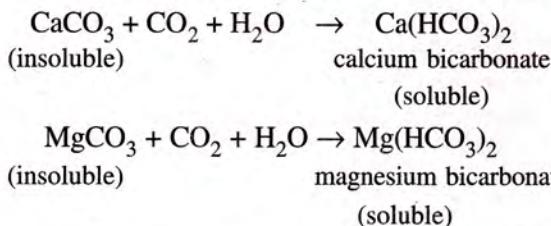
Fig. 3.14 Soft and hard water

The conical pillar which grows downwards from the roof is called **Stalacite** and the one which grows upward from the floor of the cave is called **stalagmite** (Fig. 3.15).



*Fig. 3.15 Stalactites and stalagmites*

They meet after a time. In a year some grow less than even a centimeter but some as tall as 100 cm.

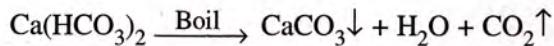


If water flows over beds of gypsum ( $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ ), a little bit of gypsum gets dissolved in water and makes it hard.

### **3.13 TYPES OF HARDNESS TEMPORARY AND PERMANENT HARDNESS**

*Hardness of water is of two types : temporary hardness and permanent hardness. Water that contains only hydrogen carbonates of calcium and magnesium is called temporary hard water.*

It is called temporary hard water because its hardness can be removed just by boiling.



The calcium ions and magnesium ions are thus removed from the water by the formation of insoluble calcium carbonate and hence the water becomes soft.

Water containing sulphates and chloride of magnesium and calcium is called permanent hard water. This hardness can not be removed by boiling.

**Note :** The removing of temporary hardness of water by boiling is, however, not a practical method; it is costly and slow.

### **3.13.1 Advantages of hard water**

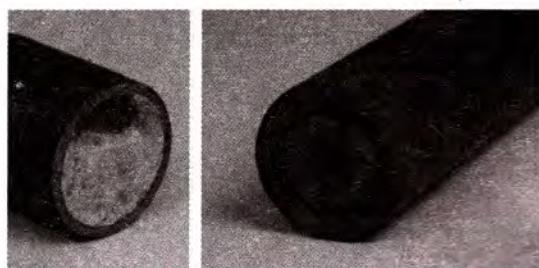
Hard water containing some dissolved salts has the following advantages.

- (i) Soft water free from dissolved salts has a very flat taste. The presence of salts in hard water makes it tasty. Hard water is therefore, used in the preparation of beverages and wines.
  - (ii) Calcium and magnesium salts present in small amounts in hard water are essential for the growth of our bones and teeth.
  - (iii) Hard water checks the poisoning of water by lead pipes. When these pipes are used for carrying water, some lead salts dissolve in water to make it poisonous calcium sulphate present in hard water forms insoluble lead sulphate in the form of a layer inside the lead pipe and this checks the lead poisoning.

### **3.13.2 Disadvantages of hard water**

**Hard water is not suitable for producing steam**

Steam is usually made in boilers which are made up of a number of narrow copper tubes surrounded by fire. As the cold water enters these tubes, it is immediately changed into steam, while the dissolved solids incapable of changing into vapour get deposited on the inner walls of the tubes. This goes on and thus making the bore of the tubes narrower and narrower (Fig. 3.16). The result is that less water flows through the tubes at one time and less steam is produced.



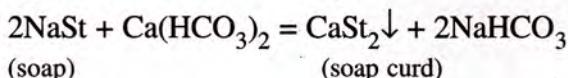
*Fig. 3.16 Boiler scales due to hard water*

*When the bore of the tube becomes very narrow, the pressure of steam increases so much that at times the boiler itself bursts. Not only that, the boiler scale formed inside the tube being a bad conductor of heat allows only part of the outside heat to go in and convert the water into steam. This results in further loss of heat. To minimise the loss, the boiler scale is removed from time to time by using a boiler compound or else water is first softened and then used. Both cost money.*

The **furring** of the tea kettle is also caused by the sediment formed on its walls from boiling hard water. This fur is nothing but carbonates of calcium and magnesium.

**Hard water is also unfit for washing purposes**

For washing purposes soap is invariably used. *Soap is chemically a sodium salt of stearic acid (an organic acid, formula  $C_{17}H_{35}COOH$ ) and has the formula  $C_{17}H_{35}COONa$ .* If the water is hard, the calcium and magnesium ions of the water combine with the negative ions of the soap to form a slimy precipitate of insoluble calcium and magnesium usually given the name of **soap-curd** or scum.



The formation of soap-curd will go on as long as there are calcium and magnesium ions present. Till then no soap lather will be formed and cleaning of cloth or body will not be possible. Moreover these precipitates are difficult to wash out from the fabrics and sometimes form rusty spots if iron salts are also present in water.

This difficulty has been solved to a great extent by using synthetic detergents.

*Synthetic detergents* such as SURF, DET, etc., are used in place of soap in domestic and laundry work. These are made by sulphonating some higher alkenes (organic compounds) with sulphuric acid and then converting them into their sodium salts by caustic soda. These sodium salts are called the detergents.

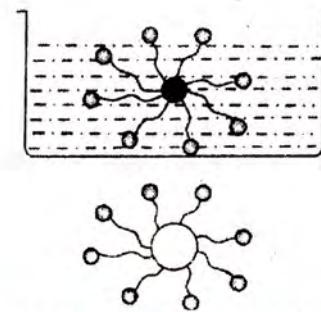
Detergents are more soluble in water than soap and are unaffected by hardness of water as their calcium salts are soluble in water.

## **Cleansing action of soaps and detergents**

When a soap or detergents is dissolved in water, the molecules gather as clusters called micelles. The tails stick inwards and the head outwards.

In cleansing, the hydrocarbon tail attaches itself to oily dirt. When water is stirred, the oily dirt tends to lift off from the dirty surface and dissociates into fragments. This gives an opportunity to other tails to stick to oil. The solution now contains small globules of oil surrounded by detergent molecules. The negatively charged heads present in water prevent the

small globules from coming together and form aggregates. Thus, the oily dirt is removed from the object.



*Fig. 3.17 Cleaning action of soap*

### 3.13.3 Removal of hardness

#### **Removal of temporary hardness :**

(1) By boiling, carbon dioxide is driven off and the soluble hydrogen carbonates are converted into insoluble carbonates and could be removed by filtration or decantation.

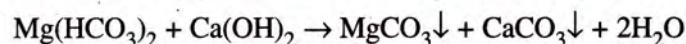
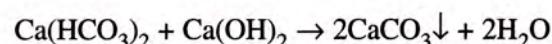


Calcium carbonate and magnesium carbonate are precipitated leaving the water soft.

This is, however, not very practical when large quantities of water are concerned.

#### (2) By addition of lime (Clark's Process).

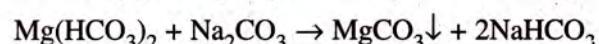
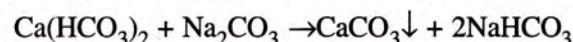
A more practical method is to add a calculated quantity of slaked lime. The following reaction takes place.



Lime is first thoroughly mixed with water in a tank and then fed into another tank containing the hard water. Revolving paddles thoroughly mix the two solutions. Most of the calcium carbonate settles down. If there is any solid left over it is removed by a filter. The process goes by the name of Clarke's process.

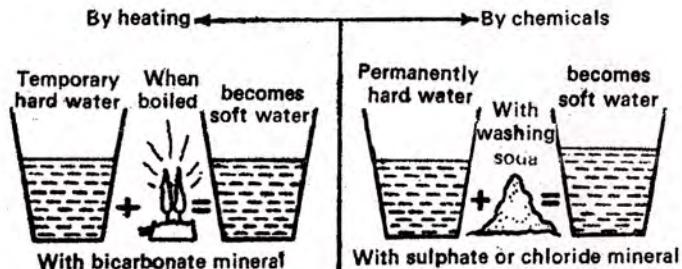
### **(3) By the addition of washing soda.**

When washing soda or soda ash is added to hard water, the corresponding insoluble carbonates settle down and can be removed by filtration.

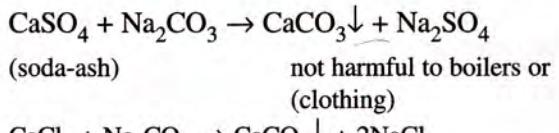


## **Removal of permanent hardness**

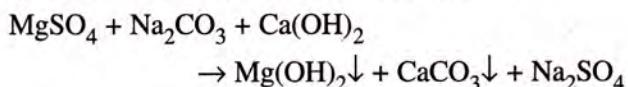
To remove permanent hardness of water, it is treated with a calculated quantity of soda ash, when the following reaction takes place



*Fig. 3.18 Softening of hard water*



If, however, some magnesium sulphate is also present in the water, lime has also to be used besides soda ash, as magnesium carbonate is not fully precipitated. In that case the reaction will be



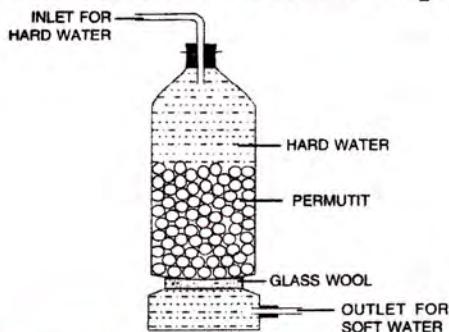
Magnesium hydroxide is less soluble than magnesium carbonate.

**Note :** The above methods are suitable, for domestic purposes only when small amounts of water have to be treated.

**If hardness is both temporary and permanent.**  
In such a case, particularly when large amounts of water are concerned, the permutit process is used.

Permutit is an artificial zeolite. Chemically, it is hydrated sodium aluminium orthosilicate, having the

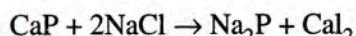
formula  $\text{Na}_2\text{Al}_2\text{Si}_2\text{O}_8 \cdot \text{XH}_2\text{O}$ . For the sake of convenience let us give it the formula  $\text{Na}_2\text{P}$ .



*Fig. 3.19 Softening of hard water by permit it*

A tall cylinder is loosely filled with lumps of permutit. When hard water containing calcium and magnesium ions percolates through these lumps, exchange of ions takes place. The sodium permutit is slowly changed into calcium and magnesium permutit, and with the removal of calcium and magnesium ions, the water becomes soft.

When no longer active, the permutit is regenerated by running a concentrated solution of brine over it and removing the calcium chloride formed by repeated washing.



### **Softening hard water, using ion exchange resins**

For laundries and boilers, water softened by the permutit process is quite satisfactory.

For scientific purposes even the presence of +ve sodium ion and -ve bicarbonate and sulphate ions is not desirable. To remove these ions, the water is passed through synthetically prepared cation (+ve ion) and anion (-ve ion) exchange resins one by one. The sodium ions are replaced by  $H^+$  ions and  $HCO_3^-$  and  $SO_4^{2-}$  ions by  $OH^-$  ions and the water gets completely deionised.

### **EXERCISE 3(C)**