COMPOUNDS OF MAGNESIUM AND CALCIUM

Compounds of Magnesium (Mg)

Magnesium salts are naturally present in seawater. The crystalline structure of magnesium is characterized by a closed-packed hexagonal lattice. Within the human body, magnesium ions play a crucial role as an essential element, and they serve as a vital component of chlorophyll, the green pigment responsible for photosynthesis in plants.

Importance ores of magnesium are

(i) Magnesite, $MgCO_3$ (ii) Dolomite, $MaCO_3 . CaCO_3$

(iii) Carnallite, $KCl, MgCl, 6H_2O$ (iv) Kieserite, $MgSO_4 . H_2O$

Seawater additionally harbors magnesium salts, with its magnesium content measuring around 0.13%. Magnesium can be acquired through the process of electrolysis using a molten mixture of anhydrous substances, which may be derived from either seawater or carnallite. Anhydrous magnesium is produced from the hydrated form by subjecting it to a flow of dry HCl gas, as direct heating leads to the formation of refractory MgO and is not a suitable method for anhydrous magnesium preparation.

$$MgCl_2 . 6H_2O \xrightarrow{\triangle} Mgo + 2HCl + 5H_2O$$

During electrolysis, the following reactions take place.

$$MgCl_2 ?\, Mg^{2+} + 2Cl^-$$

At cathode: $Mg^{2+} + 2e^- \rightarrow Mg$

At anode: $2Cl^- \rightarrow Cl_2 + 2e^-$

Magnesium Oxide (MgO)

Preparation

(i) Combustion of magnesium ribbon in air (oxygen) yields magnesium oxide (MgO) according to the reaction:

$$2Mg + O_2 \rightarrow 2MgO$$

(ii) Strongly heating magnesium hydroxide results in the formation of magnesium oxide:

$$Mg(OH)_2 \xrightarrow{\triangle} MgO + H_2O$$

(iii) The strong heating of magnesium carbonate produces magnesium oxide:

$$MgCO_3 \xrightarrow{\triangle} MgO + CO_2$$

(iv) The strong heating of magnesium nitrate leads to the production of magnesium oxide:

$$Mg(NO_3)_2 \xrightarrow{\triangle} 2MgO + 4NO_2 + O_2$$

Properties

(i) Magnesium oxide manifests as a white powder.

(ii) Functioning as a basic oxide, magnesium oxide is soluble in water.

(iii) Collectively known as magnesia cement or Sorel's cement (i.e., $MgCl_2 \cdot 5MgO \cdot H_2O$), magnesium oxide and magnesium chloride find application in dental fillings.

(iv) Reacting magnesium oxide with hydrochloric acid (HCl) or sulfuric acid (H_2SO_4) produces the respective salts:

$$MgO + 2HCl \rightarrow MgCl_2 + H_2O$$

$$MgO + H_2SO_4 \rightarrow MgSO_4 + H_2O$$

Magnesium Hydroxide ($Mg(OH)_2$)

Preparation:

(i) Magnesium hydroxide can be obtained by reacting magnesium chloride with calcium hydroxide or sodium hydroxide:

$$MgCl_2 + Ca(OH)_2 \rightarrow Mg(OH)_2 + CaCl_2$$

$$MgCl_2 + NaOH \rightarrow Mg(OH)_2 + 2NaCl$$

(ii) Magnesium oxide, when in contact with water, undergoes a reaction to yield magnesium hydroxide:

$$MgO + H_2O \rightarrow Mg(OH)_2$$

Properties:

(i) Magnesium hydroxide is characterized by its white powder appearance.

(ii) Possessing basic properties, magnesium hydroxide exhibits alkalinity.

(iii) Decomposition of magnesium hydroxide results in the formation of magnesium oxide.

(iv) Reacting magnesium hydroxide with ammonium chloride produces magnesium chloride:

$$Mg\,(OH)_2 + 2NH_4Cl \rightarrow MgCl_2 + 2NH_4OH$$

Magnesium Carbonate ($MgCO_3$)

Preparation:

(i) Magnesium carbonate is produced by introducing sodium bicarbonate into a heated solution of magnesium salt:

$$MgSO_4 + 2NaHCO_3 \rightarrow MgCO_3 + Na_2SO_4 + H_2O + CO_2$$

(ii) When magnesium sulfate reacts with sodium carbonate, a white precipitate of basic carbonate is formed. Upon suspension in water and the passage of carbon dioxide, magnesium bicarbonate is generated, and subsequent boiling results in the formation of magnesium carbonate:

$$2MgSO_4 + 2Na_2CO_3 + H_2O \rightarrow MgCO_3{\cdot}Mg(OH)_2 + 2Na_2SO_4 + CO_2$$

$$MgCO_3{\cdot}Mg(OH)_2 + 3CO_2 + H_2O \rightarrow 2Mg(HCO_3)_2$$

$$Mg(HCO_3)_2 \rightarrow MgCO_3 + CO_2 + H_2O$$

Magnesium Chloride ($MgCl_2$)

Preparation:

(i) The production of magnesium chloride involves dissolving magnesium (Mg), magnesium oxide (MgO), magnesium hydroxide ($Mg(OH)_2$), or magnesium carbonate ($MgCO_3$) in diluted hydrochloric acid (HCl):

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

$$MgO + 2HCl \rightarrow MgCl_2 + H_2O$$

$$Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + 2H_2O$$

$$MgCO_3 + 2HCl \rightarrow MgCl_2 + H_2O + CO_2$$

(ii) The fusion and subsequent cooling of carnallite ($KCl{\cdot}MgCl_2{\cdot}6H_2O$) at 176°C result in the deposition of KCl while $MgCl_2{\cdot}6H_2O$ is obtained in a fused state.

Properties:

(i) Magnesium chloride is characterized as a colorless crystalline solid, exhibiting high solubility in water.

(ii) When $MgCl_2{\cdot}6H_2O$ reacts with sodium bicarbonate, it produces magnesium carbonate. Further reaction with sulfuric acid (H_2SO_4) yields magnesium sulfate:

$$MgCl_2\ 6H_2O \xrightarrow{NaHCO_3} MgCO_3 \xrightarrow{H_2SO_4} MgSO_4\ 7H_2O$$

Magnesium Sulphate ($MgSO_4$)

Preparation:

(i) The reaction between magnesite ($MgCO_3$) and dilute sulfuric acid results in the formation of magnesium sulfate:

$$MgCO_3 + H_2SO_4 \rightarrow MgSO_4 + H_2O + CO_2$$

(ii) When dolomite ($MgCO_3{\cdot}CaCO_3$) is subjected to treatment with dilute sulfuric acid, magnesium sulfate is produced:

$$MgCO_3\ CaCO_3 + 2H_2SO_4 \rightarrow MgSO_4 + CaSO_4 + 2CO_2 + 2H_2O$$

Properties:

(i) Magnesium sulfate is a colorless crystalline compound that readily dissolves in water.

(ii) Magnesium sulfate has the ability to form double salts with alkali metal sulfates, as exemplified by the compound $K_2SO_4{\cdot}MgSO_4{\cdot}6H_2O$.

(iii) Upon intense heating, magnesium sulfate undergoes decomposition, yielding magnesium oxide (MgO).

$$MgSO_47H_2O \xrightarrow{150^0C} MgSO_4H_2O \xrightarrow{200^0C} MgSO_4 \rightarrow MgO + SO_2 + \frac{1}{2}O_2$$

Calcium and its Compounds

Calcium oxide Quick lime, Burnt lime (CaO)

Preparation

It is prepared by heating lime stone ($CaCO_3$) at 800°C

$$CaCO_3 \rightleftharpoons CaO + CO_2 \qquad \triangle H = 179.9KJ$$

Condition for Good Yield

(i) Given the reversibility of the reaction, it is imperative to promptly eliminate the carbon dioxide as it is generated to ensure the reaction proceeds in the desired direction.

(ii) At an elevated temperature, CaO will engage in a reaction with silica (SiO_2) to produce $CaSiO_3$.

Properties

(i) Action of water

$$\underset{\text{(quick lime)}}{CaO + H_2O} \longrightarrow \underset{\text{(Slaked lime)}}{Ca(OH)_2} + 15000\ \text{Caloric}$$

Paste of lime in water is called milk of lime, while its hydrate is known as lime water.

(ii) With moist chlorine:

$$CaO + Cl_2 \rightarrow CaOCl_2(\text{Bleachingpowder})$$

(iii) When heated with carbon, it forms calcium carbide.

$$CaO + 3C \xrightarrow{2000^0C} CaC_2 + CO$$

Uses

(i) In purification of sugar

(ii) As basic lining in furnaces.

Calcium hydroxide, Slaked lime $Ca(OH)_2$

Preparation

By the action of water on quick lime.

$$CaO + H_2O \rightarrow Ca(OH)_2 + \text{Heat}$$

Means it is an exothermic reaction.

Properties

(i) It dissolves only to a limited extent in water, and its solubility decreases as the temperature rises.

(ii) Action of CO_2: Lime water turns milky on passing CO_2 gas

$$Ca(OH)_2 + CO_2 \rightarrow \underset{\text{(Milkines)}}{CaCO_3} + H_2O$$

(iii) Action of chlorine

(a) Cold Condition: $2Ca(OH)_2 + 2Cl_2 \xrightarrow{\text{Cold}} CaCl_2 + \underset{\text{(Cal.Hypochlorit)}}{Ca(OCl)_2} + 2H_2O$

(b) Below 35°: $3Ca(OH)_2 + 2Cl_2 \xrightarrow[35^0 C]{Below} CaOCl_2 + H_2O$

(c) On gentle Heating: $6Ca(OH)_2 + 6cl_2 \xrightarrow{Heat} 5CaOCl_2 + H_2O$

(d) On Red Hot: $2Ca(OH)_2 + 2Cl_2 \xrightarrow{Re\ dHeat} 2CaCl_2 + Ca(ClO_3) + H_2O$

(iv) Action of Ammonia: $Ca(OH)_2 + 2NH_4Cl \rightarrow CaCl_2 + 2NH_3 + 6H_2O$

Uses

(i) To make hard water softer.

(ii) Used in the refining process of sugar and coal gas.

(iii) In the production of bleaching powder, white wash, plaster, and similar products.

Calcium Carbonate, $CaCO_3$

Calcium carbonate occurs naturally in various forms such as limestone, marble, coral, Iceland spar, calcite, chalk, dolomite, and more.

Preparation

This substance is produced as a white powder, commonly referred to as precipitated chalk, by dissolving marble or limestone in hydrochloric acid, and then causing precipitation with sodium or ammonium carbonate.

$$CaCO_3 2HCl \rightarrow CaCl_2 + H_2O + CO_2$$
$$CaCl_2(NH_4)_2CO_3 \rightarrow CaCO_3 + 2NH_4CL$$

In the laboratory, it is prepared by passing CO_2 through lime-water or by adding sod. or ammo carbonate solution of $CaCl_2$.

$$Ca(OH)_2 + CO, \rightarrow CaCO_3 \downarrow + H_2O$$

Properties

(i) This substance presents as a white powder and is found in two crystalline structures: calcite and aragonite.

(ii) It is insoluble in water but dissolves in the presence of CO_2 due to the formation of Calcium bicarbonate.

$$CaCO_3 + H_2O + CO_2 \rightarrow Ca(HCO_3)_2$$

Calcium Bicarbonate, $(Ca(HCO_3)_2)$

It is not present in a solid state, but its solution can be created by introducing carbon dioxide (CO_2) gas into a water-based suspension of calcium carbonate. Upon heating, this solution breaks down into calcium carbonate while releasing carbon dioxide gas.

Calcium Sulphate,($CsSO_4$)

It occurs in nature as anhydride $(CaSO_4)$ and gypsum $(CaSO_4 \cdot 2H_2O)$

Preparation

It may be prepared by adding dilute sulphuric acid or soluble sulphate (e.g., Na_2SO_4) to the solution of a calcium salt.

$$CaCl_2 + H_2SO_4 \rightarrow CaSO_4 + 2HCl$$

Properties

(i) This substance is in the form of a white crystalline solid. Similar to slaked lime, it displays limited solubility in cold water, and this solubility decreases even further as the temperature rises.

(ii) It readily dissolves in dilute acids, including acetic acid, at a faster rate than it does in water. Additionally, it can dissolve in an ammonium sulfate solution, resulting in the formation of calcium sulfate ammonium sulfate dihydrate, known as:

$$CaSO_4(NH_4)SO_4H_2O$$

(iii) When monoclinic gypsum crystals are subjected to heat, they initially undergo a transformation into an orthorhombic structure without any loss of water. Upon further heating to 120°C, about three-fourths of their crystallization water is released, leading to the formation of hemihydrate, which can be represented as:

$$(2CaSO_4) + H_2O \text{ or } CaSO_4\frac{1}{2}H_2O + 1\frac{1}{2}H_2O$$

Plaster of Paris, which is a white powder, can be converted back into its dihydrate form when mixed with water, as it reabsorbs the water of crystallization. This process results in the substance solidifying into a hard mass with a slight expansion. The setting of plaster of Paris typically takes approximately 10–15 minutes and can be either accelerated by adding common salt or delayed by the use of alum or borax.

$$\underset{\text{Plaster of paris}}{CaSO_4.\frac{1}{2}H_2O} \xrightarrow[\text{setting}]{H_2O} \underset{\text{Orthor hom bic}}{CaSO_4.2H_2O} \xrightarrow{\text{hardening}} \underset{\text{Monochlinic (gypsum)}}{CaSO_4.2H_2O}$$

Plaster of Paris of gypsum when heated to about 200°C is converted into anhydrous calcium sulphate. The anhydrous from (anhydrite) is known as dead burnt because it does not set like Plaster of Paris when moistened with water.

Calcium Chloride ($CaCl_2$)

Preparation

When a solution containing lime or calcium carbonate in hydrochloric acid (HCl) is evaporated, it precipitates as hygroscopic crystals.

$$CaCO_3 + 2HCL \rightarrow CaCl_2 + H_2CO_3$$

But is separates out from the reaction mixture as $CaCl_2.6H_2O$. The anhydrous salt is obtained on heating above 200°C.

Properties

This is a transparent, hygroscopic compound that readily dissolves in water. The dehydrated form of the compound is an effective desiccant.

Calcium Carbide, CaC_2

Calcium carbide is produced through the heating of quicklime and coke in an electric furnace.

$$CaO + 3C \rightarrow CaC_2 + CO\uparrow$$

CaC_2 on reaction with water produces acetylene gas

$$CaC_2 + 2H_2O \rightarrow Ca(OH)_2 + C_2H_2$$

Calcium Hydroxide Solubility in Water

Calcium Hydroxide

Preparation:

(i) Obtained from quicklime: The reaction involves the combination of calcium oxide (CaO) with water (H_2O) to produce calcium hydroxide ($Ca(OH)_2$) according to the equation:

$$CaO + H_2O \rightarrow Ca(OH)_2.$$

(ii) Derived from calcium chloride ($CaCl_2$): The process entails the reaction between calcium chloride ($CaCl_2$) and sodium hydroxide (NaOH), resulting in the formation of calcium hydroxide ($Ca(OH)_2$) and sodium chloride (NaCl). The balanced chemical equation for this reaction is:

$$CaCl_2 + 2NaOH \rightarrow Ca(OH)_2 + 2NaCl.$$

$$\begin{tikzcd}
\underset{(CaO)}{\text{Quick lime}} \arrow[r, "H_2O"] & \underset{Ca(OH)_2}{\text{Slaked lime}} \arrow[d] \\
\underset{Ca\ (OH)_2}{\text{Suspension Milk of lime}} & \text{Suspended in water} \arrow[l] \arrow[d, "\text{Filter}"] \\
 & \underset{Ca\ (OH)_2}{\text{Clear solution (lime water)}}
\end{tikzcd}$$

Properties

When exposed to carbon dioxide (CO_2), it produces calcium carbonate ($CaCO_3$) and calcium bicarbonate ($Ca(HCO_3)_2$.

$$\underset{\text{Lme water}}{Ca(OH)_2} + CO_2 \rightarrow \underset{\text{miay}}{CaCO_3} \downarrow + H_2O$$

Upon extended exposure to CO_2, the cloudiness vanishes as a result of the creation of calcium bicarbonate ($Ca(HCO_3)_2$.

$$CaCO_3 + H_2O + CO_2 \rightarrow \underset{\text{Soluble}}{Ca(HCO_3)}$$

Solubility in Water

(i) Order of solubility in sulfates:
Beryllium sulfate ($BeSO_4$) > Magnesium sulfate ($MgSO_4$) > Calcium sulfate ($CaSO_4$) > Strontium sulfate ($SrSO_4$) > Barium sulfate ($BaSO_4$).

(ii) Order of solubility in hydroxides:
Beryllium hydroxide ($Be(OH)_2$) < Magnesium hydroxide ($Mg(OH)_2$) < Calcium hydroxide ($Ca(OH)_2$) < Strontium hydroxide ($Sr(OH)_2$) < Barium hydroxide ($Ba(OH)_2$).

(iii) Order of solubility in fluorides:
Beryllium fluoride (BeF_2) > Barium fluoride (BaF_2) > Strontium fluoride (SrF_2) > Magnesium fluoride (MgF_2) > Calcium fluoride (CaF_2).

(iv) Order of solubility in chlorides:
Beryllium chloride ($BeCl_2$) > Calcium chloride ($CaCl_2$) > Magnesium chloride ($MgCl_2$) > Strontium chloride ($SrCl_2$) > Barium chloride ($BaCl_2$).

(v) Order of solubility in bromides:
Beryllium bromide ($BeBr_2$) > Calcium bromide ($CaBr_2$) > Magnesium bromide ($MgBr_2$) > Barium bromide ($BaBr_2$) > Strontium bromide ($SrBr_2$).

(vi) Stability of iodides:
Calcium iodide (Cal_2) > Barium iodide (Bal_2) > Strontium iodide (Srl_2) > Magnesium iodide (Mgl_2)
[Note: Beryllium iodide (Bel_2) decomposes upon dissolution].

(vii) Order of solubility in nitrates:
Calcium nitrate ($Ca(NO_3)_2$) > Beryllium nitrate ($Be(NO_3)_2$) > Strontium nitrate ($Sr(NO_3)_2$) > Magnesium nitrate ($Mg(NO_3)_2$) > Barium nitrate ($Ba(NO_3)_2$).