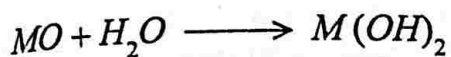


## General characteristics of compounds

### 1. Oxides and hydroxides

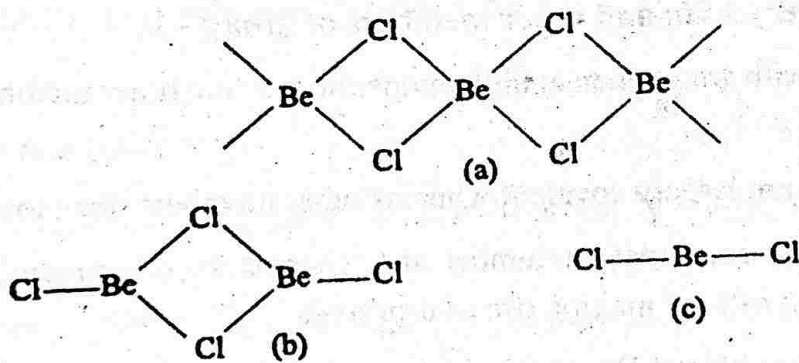
(i) **Oxides and hydroxides:** Alkaline earth metals burn in air or oxygen to form their oxides.  $BeO$  is covalent while all other oxides are ionic.

$BeO$  is amphoteric in character, while the oxides of the rest of the elements in group 2 are basic. The oxides of these metals (except  $BeO$ ) react with water to form their corresponding hydroxides.



The hydroxides of alkaline earth metals are basic except  $Be(OH)_2$  which is amphoteric. The basic strength increases from  $Mg(OH)_2$  to  $Ba(OH)_2$ . The solubility and thermal stability of hydroxides also increase downward in the group.  $Be(OH)_2$  and  $Mg(OH)_2$  are almost insoluble,  $Ca(OH)_2$  is sparingly soluble, while  $Sr(OH)_2$  and  $Ba(OH)_2$  are increasingly more soluble.

(ii) **Halides:** Beryllium halides are covalent. Other halides of these metals are ionic. Beryllium chloride is soluble in organic solvents indicating its covalent nature. In the solid state  $BeCl_2$  has a polymeric chain structure involving  $Be-Cl-Be$  bridge



(a) structure of solid  $BeCl_2$

(b)  $BeCl_2$  dimer

(c) monomer

In the vapour phase  $BeCl_2$  tends to form a chloro-bridged dimer which dissociates into the linear monomer at elevated temperatures of the order of 1200 K.

Halides of the other group 2 metals (except fluorides) are soluble in water. The tendency to form halide hydrates decreases down the group. e.g.,  $MgCl_2 \cdot 8H_2O$ ,  $CaCl_2 \cdot 6H_2O$ ,  $SrCl_2 \cdot 6H_2O$ ,  $BaCl_2 \cdot 2H_2O$ .

(iii) **Salts of oxoacids (Sulphates and carbonates):** The solubility of sulphates and carbonates of alkaline earth metals decreases down the group with increase in atomic number. This is in contrast to the behaviour of their hydroxides. The difference in solubilities can be understood from the energy of hydration and lattice energy values. As the energy of hydration decreases down the group, the solubilities of sulphates and carbonates of group 2 elements decrease in the order

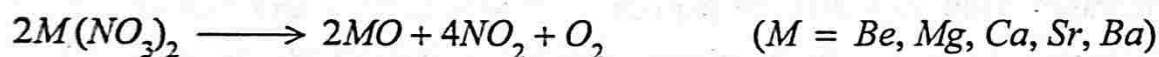


Carbonate of alkaline earth metals are insoluble in water.

While the sulphates are stable towards heat, the carbonates decompose on heating to form the oxide. The thermal stability of carbonates increases down the group.

(iv) **Nitrates:** The nitrates are prepared by treating the carbonate with dilute  $HNO_3$ .

All nitrates of alkaline earth metals decompose on heating to form the oxide, nitrogen dioxide and oxygen.



### Anomalous behaviour of beryllium

Beryllium differs from the rest of the elements in some of its properties. This is due to (i) the small size of *Be* atom (ii) relatively high ionisation enthalpy and (iii) the absence of vacant *d*-orbitals.

### Difference between beryllium and other members of group - 2

1. *Be* does not react with water even at high temperature while other members of the group react with water.
2. Compounds of *Be* are largely covalent whereas other members form ionic compounds.
3. *Be* shows maximum coordination number of 4 whereas the other members of the group can have a coordination of 6 by making use of *d*-orbitals
4. The oxide and hydroxides of *Be* are amphoteric whereas those of other members are basic.

### Diagonal relationship between *Be* and *Al*

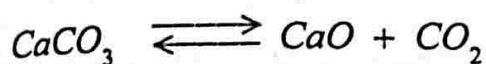
1. The hydroxides of *Be* and *Al* are amphoteric.  $Be(OH)_2$  dissolves in excess alkali to give beryllate ion  $[Be(OH)_4]^{2-}$  just like  $Al(OH)_3$  which gives aluminate ion  $[Al(OH)_4]^-$
2. Like *Al*, *Be* is not readily attacked by acids because of the presence of an oxide film on its surface.
3. Both  $BeCl_2$  and  $AlCl_3$  are soluble in organic solvents and are strong Lewis acids.

4. Both beryllium chloride and aluminium chloride have bridged chloride structures in the vapour phase.
5. Beryllium and aluminium ions have strong tendency to form complexes.

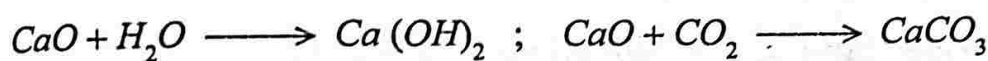
## Compounds of Calcium

### Calcium oxide (CaO)

Calcium oxide, known as *quick lime* is prepared on a large scale by heating limestone in a rotary kiln at 1070-1270K.



It is a white amorphous solid which melts at 2870K. On exposure to atmosphere, it absorbs moisture and carbon dioxide.



Since CaO is a basic oxide, it combines with acidic oxides on heating



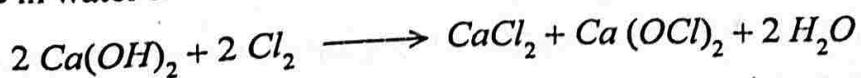
### Uses

(i) Quick lime is used in building industry. (ii) In the purification of sugar and in the manufacture of dye stuffs. (iii) It is used as a flux in metallurgy. (iv) Used in the manufacture of  $\text{Na}_2\text{CO}_3$ .

### Calcium hydroxide (Ca(OH)<sub>2</sub>)

Calcium hydroxide, known as *slaked lime* is obtained by adding limited amount of water to quick lime. Addition of small amount of water to quick lime produces large amount of heat and the process is called *slaking of lime*.  $\text{CaO} + \text{H}_2\text{O} \longrightarrow \text{Ca(OH)}_2$

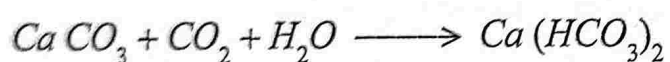
Slaked lime is a white amorphous solid sparingly soluble in water. A suspension of slaked lime in water is called *milk of lime*. Chlorine reacts with slaked lime to form bleaching powder.



An aqueous solution of slaked lime is known as lime water. When carbon dioxide gas is passed through lime water, it turns milky due to the formation of insoluble  $\text{CaCO}_3$ .



When excess  $\text{CO}_2$  is passed, the milkiness disappears due to the formation of soluble calcium bicarbonate.



#### Uses

(i) Used in white wash due to its disinfectant nature. (ii) Slaked lime is used for softening hard water. (iii) It is used as a building material and for absorbing acidic gases.