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s-Block Elements

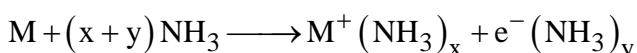
	Characteristic	Group 1 (Alkali metals)	Group 2 (Alkaline earth metals)
(A)	General Introduction		
1.	Elements present	^3Li , ^{11}Na , ^{19}K , ^{37}Rb , ^{55}Cs and ^{87}Fr	^4Be , ^{12}Mg , ^{20}Ca , ^{38}Sr , ^{56}Ba , ^{88}Ra
2.	Name	They are known as alkali metals because their oxides and hydroxides when treated with water form alkaline solutions	Oxides of calcium, barium and strontium were known before these metals were discovered and since these oxides were alkaline and found in the earth, the metals were named as alkaline earth metals. Now-a-days this term is extended to all elements of group 2 or II-A
3.	Electronic configuration	[Noble gas]ns ¹	[Noble gas] ns ²
(B)	Physical Properties		
1.	Atomic radii	Largest in their respective periods. Down the group, the atomic radius increases .	Large but smaller than the corresponding alkali metals. Down the group, the atomic radius increases .
2.	Ionic Radii	Ions are formed by the loss of one electron Hence ionic radius is smaller than atomic radius but trend remains the same as above i.e. increases down the group.	Ions are formed by the loss of two electrons. The trend remains the same as above i.e. increases down the group.
3.	Ionization energy	Lowest in the their respective periods. Down the group it decreases	They have fairly low ionization energies but greater than the corresponding elements of Group 1. These decrease down the group.
4.	Oxidation state	+1 because they lose one electron to acquire stable noble gas configuration.	+2 because they lose two electrons to acquire stable noble gas configuration.
5.	Metallic character	These are typical soft metals and can be easily cut with a knife. Because of low ionization energies, they are highly electropositive. The electropositive character increases down the group.	These are also metals. With respect to alkali metals they are less metallic and less electropositive. Their metallic character increases down the group.

	Characteristic	Group 1 (Alkali metals)	Group 2 (Alkaline earth metals)																		
6.	Melting points and Boiling points	Their melting point and boiling point are low and decrease down the group because of large size and weak intermetallic bonds.	Their melting point and boiling point are higher than alkali metals because of smaller size and more closely packed crystal lattice. There is no regular trend down the group.																		
7.	Nature of the bonds formed	They form ionic bonds. The tendency to form ionic bond increases down the group.	Because of small size and high I.E., Be has a tendency to form covalent compounds. Mg also shows some tendency for covalency. All other elements form ionic compounds.																		
8.	Density	Their densities are quite low due to their large size. Li, Na and K are lighter than water. The densities increases from Li to Cs although density of potassium is lower than that of sodium.	These are denser than alkali metals because of smaller size. Density decreases from Be to Ca and then increases from Sr to Ra.																		
9.	Flame colouration	<p>These metals or their salts impart characteristic colour to the flame. The colour deepens from Li to Cs because for the same excitation energy, the electrons jump to different energy levels.</p> <table style="margin-left: auto; margin-right: auto;"> <tr> <td style="text-align: center;">Li</td> <td style="text-align: center;">Na</td> </tr> <tr> <td style="text-align: center;">Crimson red</td> <td style="text-align: center;">Golden Yellow</td> </tr> <tr> <td style="text-align: center;">K</td> <td style="text-align: center;">Rb</td> <td style="text-align: center;">Cs</td> </tr> <tr> <td style="text-align: center;">Pale</td> <td style="text-align: center;">Violet</td> <td style="text-align: center;">Violet</td> </tr> </table>	Li	Na	Crimson red	Golden Yellow	K	Rb	Cs	Pale	Violet	Violet	<p>These metals also impart characteristic colour to the flame which deepens from Ca to Ra because of the same reasons.</p> <table style="margin-left: auto; margin-right: auto;"> <tr> <td style="text-align: center;">Ca</td> <td style="text-align: center;">Sr</td> </tr> <tr> <td style="text-align: center;">Brick red</td> <td style="text-align: center;">Crimson</td> </tr> <tr> <td style="text-align: center;">Ba</td> <td style="text-align: center;">Ra</td> </tr> <tr> <td style="text-align: center;">Apple green</td> <td style="text-align: center;">Crimson</td> </tr> </table>	Ca	Sr	Brick red	Crimson	Ba	Ra	Apple green	Crimson
Li	Na																				
Crimson red	Golden Yellow																				
K	Rb	Cs																			
Pale	Violet	Violet																			
Ca	Sr																				
Brick red	Crimson																				
Ba	Ra																				
Apple green	Crimson																				
(C)	Chemical Properties																				
1.	Reducing character	<p>These metals are strong reducing agents.</p> <p>Their reducing power increases down the group,</p> <p>Exception. Li is the strongest reducing agent although it has highest I.E. in the group.</p>	<p>These are weaker reducing agents with respect to alkali metals because of greater ionization energies. The reducing power increases down the group because of decrease in ionization energy.</p>																		

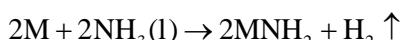
	Characteristic	Group 1 (Alkali metals)	Group 2 (Alkaline earth metals)
2.	Reaction with oxygen	Li on heating forms lithium monoxide (Li_2O), sodium forms mainly sodium peroxide (Na_2O_2); K, Rb and Cs form superoxide (MO_2).	Be, Mg and Ca form their oxides (BeO , MgO and CaO), Sr and Ba form peroxides (SrO_2 , BaO_2). This shows that the reactivity increases down the group.
3.	Reaction with water .	<p>These metals and their oxides react with water and form their hydroxides</p> $2\text{M} + 2\text{H}_2\text{O} \rightarrow 2\text{MOH} + \text{H}_2$ $\text{M}_2\text{O} + \text{H}_2\text{O} \rightarrow 2\text{MOH}$ <p>The reactivity of these metals with water increases down the group.</p> <p>(i) Basic character. All the hydroxides are basic in nature. The basic strength increases down the group.</p> <p>(ii) Solubility in water. All these are soluble in water.</p> <p>(iii) Stability. Except LiOH, all other hydroxides are thermally stable</p> $2\text{LiOH} \xrightarrow{\Delta} \text{Li}_2\text{O} + \text{H}_2\text{O}$	$\text{Be} + \text{H}_2\text{O} \rightarrow \text{No reaction}$ $\text{Mg} + \text{H}_2\text{O} \rightarrow \text{MgO} + \text{H}_2 \quad (\text{steam})$ $\text{M} + 2\text{H}_2\text{O}(\text{cold}) \rightarrow \text{M(OH)}_2 + \text{H}_2$ <p>(Ca, Sr, Ba)</p> <p>(i) Basic character. They are weaker bases than alkali metal hydroxides. The basic character increases down the group. Be(OH)_2 is amphoteric.</p> <p>(ii) Solubility in water. Less soluble. The solubility increases down the group because lattice energy decreases whereas hydration energy remains almost unchanged.</p>
4.	Reaction with hydrogen	<p>On heating with H_2, all alkali metals form ionic hydrides</p> $2\text{M} + \text{H}_2 \xrightarrow{\Delta} 2\text{M}^+ \text{H}^-$ <p>Ease of formation of hydrides decreases from Li to Cs. Li combines with H_2 at approx 800°C.</p> <p>The stability of hydrides decreases from LiH to CsH.</p> <p>Ionic nature of hydrides increases from LiH to CsH.</p> <p>Ionic hydrides are decomposed by water.</p> $\text{M}^+ \text{H}^- + \text{H}_2\text{O} \rightarrow \text{MOH(aq)} + \text{H}_2 \uparrow$ <p>Alkali metal hydrides are strong reducing agents and their reducing nature increases from LiH to CsH.</p>	<p>Except Be, all other combine with H_2 directly on heating to form the hydrides (MH_2). BeH_2 is obtained as follows:</p> $2\text{BeCl}_2 + \text{LiAlH}_4 \rightarrow 2\text{BeH}_2 + \text{LiCl} + \text{AlCl}_3$

Some Additional Characteristics of Alkali Metals

1. Photoelectric effect. Except Li (which has highest ionization energy), all metals of this group exhibit in a liquid photoelectric effect which increases down the group.
2. Solubility in liquid ammonia. All alkali metals dissolve in liquid ammonia and produce intense blue solution. This blue solution conducts electricity and possesses strong reducing power. All these properties of this solution are explained due to the presence of ammoniated (solvated) electrons.

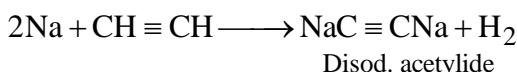
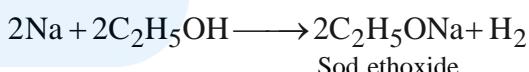


- The dilute solution of alkali metals in liquid ammonia is paramagnetic in nature.
- The paramagnetic character of the above solution decreases with increasing concentration
- The reducing power of alkali metals is enhanced in ammonia solution.
- On standing the blue solution slowly liberates hydrogen in the presence of impurities or catalyst such as Fe.

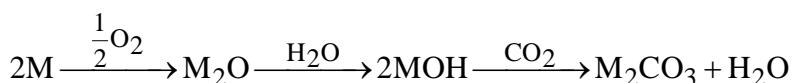


- Li forms imide with ammonia Li_2NH , while other alkali metals form amides MNH_2 .
- Highly pure dilute solution of sodium in pure liquid ammonia does not evolve hydrogen gas.
- The blue colour of the solution is due to absorption of red frequencies by ammoniated electrons.
- With increase in concentration of alkali metal in ammonia, cluster formation of metal ions takes place and finally its colour appears like metallic copper.
- The high electrical conductivity of the ammonia solution is due to ammoniated electrons and ammoniated metal ions.
- The ammonia solutions are sometimes called expanded metals because of considerable expansion of total volume.

3. Reaction with alcohol and acetylene. They liberate H_2 gas e.g.



4. Effect of air and moisture. They are converted into oxides, hydroxides and carbonates.

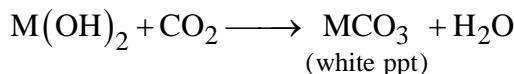


That is why they are kept under kerosene oil.

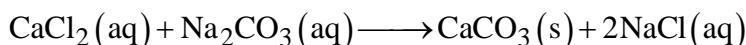
5. Hydration of ions. Alkali metals are highly hydrated in aqueous solutions. Smaller the size, greater is the hydration. Hence hydrated ionic radii are in the order : $Li^+ > Na^+ > K^+ > Rb^+ > Cs^+$. Ionic mobilities are in the opposite order.

Some Additional Characteristics of Alkaline Earth Metals and their Compounds

1. Formation of complexes. Due to smaller size and higher charge, beryllium and magnesium have the tendency to form complexes.
2. Carbonates. Preparation (i) By passing CO_2 through metal hydroxide

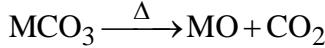


(ii) By treating the salt such as $CaCl_2$ with sodium or ammonium carbonate

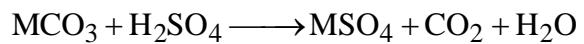
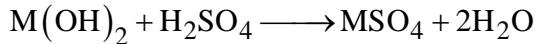
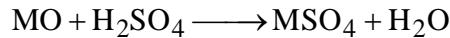
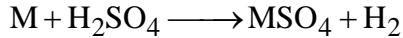


Properties

- (i) Solubility : Solubility of carbonates (from CaCO_3 to BaCO_3) decreases down the group due to decrease hydration energy
- (ii) Stability : They decompose on heating to give metal oxide and CO_2 . The temp. of decomposition increases from Be to Ba. BeCO_3 is unstable.

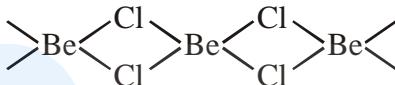


3. Sulphates. Preparation - They are obtained by action of H_2SO_4 on metals, metal oxides, metal hydroxides and metal carbonates.



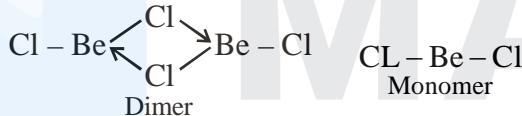
Properties.

- (i) Solubility : Solubility of sulphates (From CaSO_4 to BaSO_4) decreases down the group due to decrease of hydration energy.
 - (ii) Stability : They are thermally stable due to their high lattice energy.
4. Structure of BeCl_2 . In the solid phase; it has a polymeric structure.



The bridge Cl atoms are linked to Be atoms by three centre two electron ($3c - 2e$) bonds.

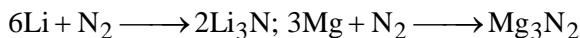
In the vapour phase, it forms dimers which dissociate to form linear monomers at 1200 K.



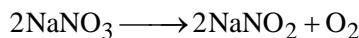
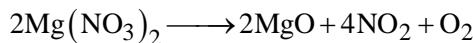
Anomalous Behaviour of Lithium and its Diagonal Relationship with Magnesium

Due to small size and high ionization energy, Li differs from the rest members of its own group and resembles diagonally with Mg as

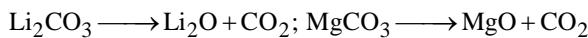
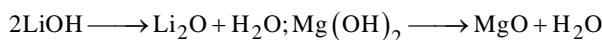
- (i) Lithium like Mg is harder while other alkali metals are soft.
- (ii) With oxygen, Li forms monoxide while others form peroxides or superoxides.
- (iii) Melting and boiling points of both Li and Mg are high with respect to other alkali metals.
- (iv) Li like Mg forms ionic nitrides while other alkali metals do not.



- (v) Nitrates of Li and Mg on heating decompose to give NO_2 and O_2 (both) while other alkali metal nitrates give O_2 .



- (vi) Hydroxides and carbonates of both Li and Mg decompose on heating while corresponding salts of other alkali metals are stable.

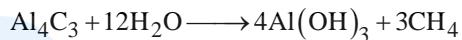
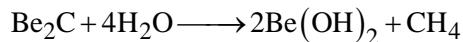


- (vii) LiOH like $Mg(OH)_2$ is a weak base while hydroxides of other alkali metals are strong bases.
- (viii) Because of covalent character, chlorides of both Li and Mg are soluble in organic solvents while those of other alkali metals are insoluble.
- (ix) Hydroxides, carbonates and fluorides of both Li and Mg are sparingly soluble in water whereas corresponding salts of other alkali metals are soluble.

Anomalous Behaviour of Beryllium and its Diagonal Relationship with Aluminium

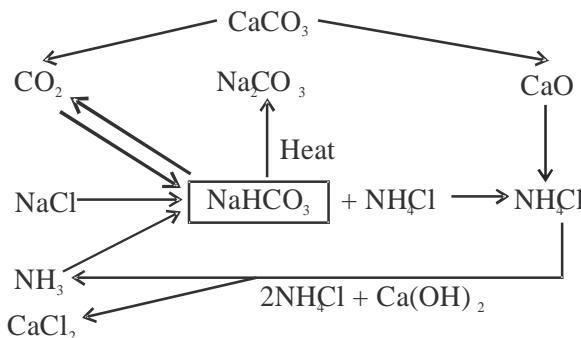
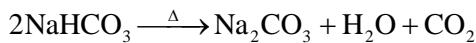
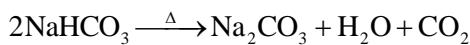
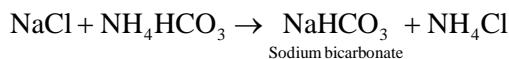
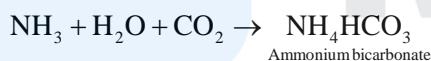
Beryllium differs from the rest members of its own group and resembles diagonally with Al in the following characteristics:

- (i) In HNO_3 , both form protective layer of oxide whereas other alkaline earth metals do not.
- (ii) Both dissolve in strong alkalis to form complexes $[Be(OH)_4]^{2-}$, $[Al(OH)_4]^-$
- (iii) Oxides of both (BeO , Al_2O_3) are hard, high melting insoluble solids, amphoteric in nature whereas oxides of other alkaline earth metals are basic.
- (iv) $BeCl_2$ is covalent with polymeric structure like aluminium chloride which has a dimeric structure, Al_2Cl_6 . The chlorides of other alkaline metals are ionic and monomeric.
- (v) Carbides of both Be and Al react with water producing methane gas whereas carbides of other alkaline earth metals give acetylene gas.



IMPORTANT COMPOUNDS

1. Sodium Carbonate or Washing Soda ($Na_2CO_3 \cdot 10H_2O$) : It is manufactured by ammonia soda process or Solvay's process :



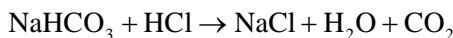
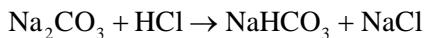
Physical Properties:

- (i) Sodium carbonate is a white crystalline solid. It is known in several hydrated forms.
- (ii) On heating, the monohydrate changes into the anhydrous form which does not decompose on further heating even to redness. It is the amorphous powder and called soda ash. It melts at 852°C.
- (iii) It is soluble in water with evolution of considerable amount of heat.



Chemical Properties

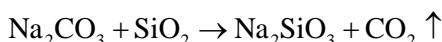
- (i) Action of acid :



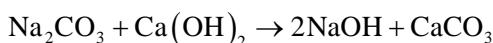
- (ii) Action of CO_2 :



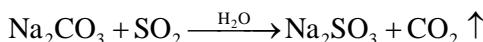
- (iii) Action of silica :



- (iv) Action of Slaked lime :



- (v) Action of sulphur dioxide:

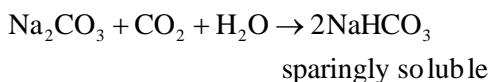


Uses: Sodium carbonate is used :

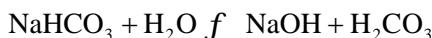
- (i) In laundries and in softening of water as washing soda
- (ii) In manufacture of glass, sodium silicate, paper, borax, soap powders, caustic soda, etc.
- (iii) in textile and petroleum refining.
- (iv) for the preparation of various carbonates of metals.

2. Sodium Bicarbonate (Baking Soda), NaHCO_3

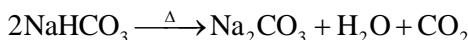
It is obtained as the intermediate product in the Solvay ammonia soda process. Normal carbonate can be changed to bicarbonate by passing carbon dioxide through its saturated solution.



Properties : It is white crystalline solid, sparingly soluble in water. The solution is alkaline in nature due to hydrolysis.



On heating, it loses carbon dioxide and water forming sodium carbonate.



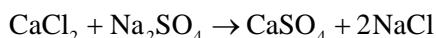
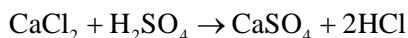
Uses:

- (i) It is used as a medicine (sodabicarb) to neutralize the acidity in the stomach.
- (ii) It is largely used for making baking powder. Baking powder is mixture of potassium hydrogen tartrate and sodium bicarbonate.
- (iii) It is used in making effervescent drinks.
- (iv) It is used in fire extinguishers.

- (v) It is used for production of carbon dioxide.
3. Calcium Sulphate

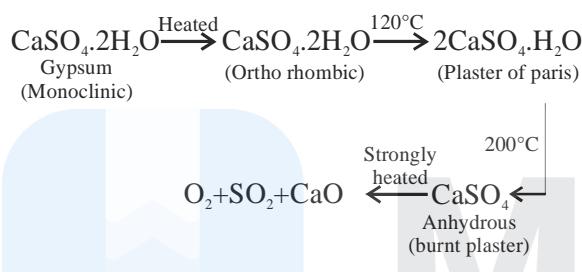
It is found in nature as anhydride (CaSO_4) and gypsum ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$).

It can be prepared by reacting any calcium salt with either sulphuric acid or a soluble sulphate.



Properties :

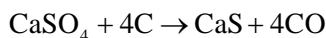
- (i) It is a white crystalline solid. It is sparingly soluble in water and solubility decreases as temperature increases.
- (ii) It dissolve in dilute acids. It also dissolves in ammonium sulphate due to the formation of double sulphate, $(\text{NH}_4)_2\text{SO}_4 \cdot \text{CaSO}_4 \cdot \text{H}_2\text{O}$.
- (iii) Effect of Temperature : Gypsum when heated first changes from monoclinic form to orthorhombic without loss of water. At 120°C , it loses three-fourth of its water of crystallisation and forms hemihydrates, $(2\text{CaSO}_4 \cdot \text{H}_2\text{O})$ which is commonly known as Plaster of Paris. At 200°C , it becomes anhydrous. The anhydrous form is known as burnt plaster or dead plaster.



- (iv) A suspension of gypsum when saturated with ammonia and carbon dioxide forms ammonium sulphate, a nitrogenous fertilizer.



- (v) When strongly heated with carbon, it forms calcium sulphide.

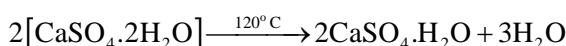


Uses : It is used

- (i) for the manufacture of Plaster of Paris, cement, ammonium sulphate, sulphuric acid.
- (ii) for preparing blackboard chalk.
- (iii) in anhydrous form as drying agent.

4. Plaster of Paris, $2\text{CaSO}_4 \cdot \text{H}_2\text{O}$ or $\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$ (Calcium sulphate hemihydrate)

Preparation : It is obtained when gypsum, calcium sulphate dihydrate ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$), is heated at 120°C .



Gypsum Plaster of Paris

(Calcium sulphate hemihydrate)

Properties :

- (i) Plaster of Paris is white powder.

(ii) It has the property of setting to a hard mass when a paste with water is allowed to stand aside for sometime. Slight expansion occurs during the setting as water is absorbed to reform $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ (gypsum). The setting process is exothermic. The process of setting takes place in two stages. In the first stage, there is conversion of Plaster of Paris into orthorhombic form of gypsum (setting step) and in the second stage orthorhombic form changes into monoclinic form (hardening step).



The setting of plaster of paris may be catalysed by sodium chloride while it is retarded by borax or alum. Addition of alum of Plaster of Paris makes the setting very hard. The mixture is known as Keene's cement.

(iii) When plaster of Paris is heated at 200°C , it forms anhydrous calcium sulphate which is known as dead plaster. It has no setting property as it takes up water only very slowly.

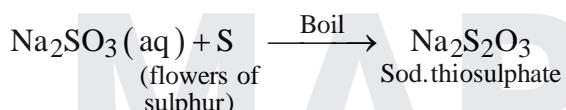
Uses : It is used

- (i) in surgery for setting broken or dislocated bones.
 - (ii) in making casts for statues,toys, surgical instruments, etc.
 - (iii) in making casts in dentistry.
 - (iv) in making blackboard chalk.

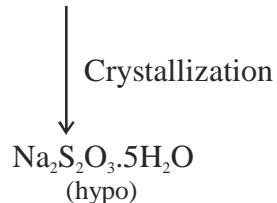
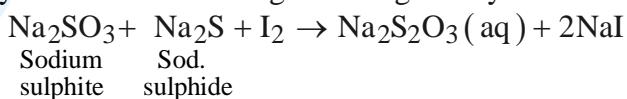
5. Sodium Thiosulphate, Hypo, $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$:

Preparation :

(i) By boiling solution of sodium sulphite with flowers of sulphur till the alkaline reaction has disappeared.



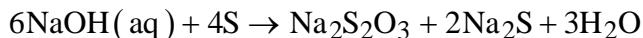
(ii) By Spring's method : A mixture of solution of Na_2S and Na_2SO_3 is treated with a calculated quantity of iodine. Fractional crystallization of resulting solution gives crystals of $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$



(iii) By passing a current of SO_2 gas in a solution of sodium carbonate and sodium sulphide.



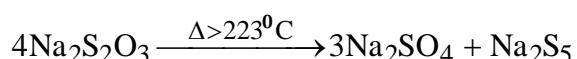
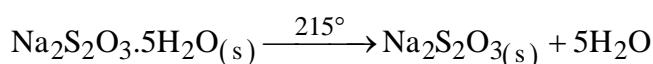
(iv) By boiling sulphur with NaOH solution :



Properties :

(i) It is colourless, crystalline, efflorescent solid. It melts at 48°C. It is extremely soluble in water and has a tendency to form supersaturated solution.

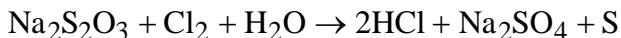
(ii) Effect of heat.



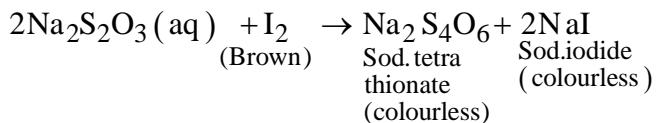
(iii) Action of dilute acids : Sulphur is precipitated and SO_2 gas is evolved.



(iv) Action with halogens : Chlorine oxidises it to sodium sulphate and sulphur is precipitated.



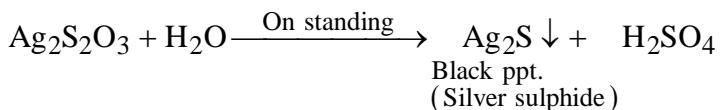
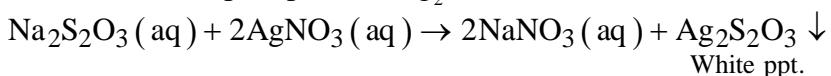
Iodine oxidises sodium thiosulphate to sodium tetrathionate.



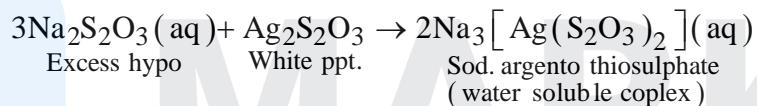
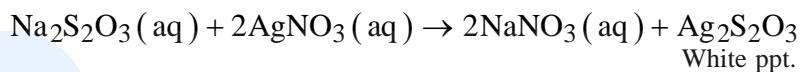
- This reaction is used in iodoetric titrations in volvetry. The indicator of the titration between iodine and hypo is starch.

(v) Action of silver nitrate solution :

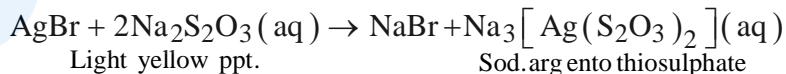
(a) When AgNO_3 is in excess : A black precipitate of Ag_2S is obtained.



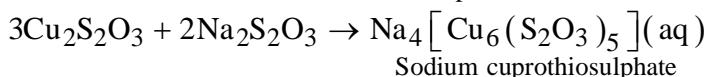
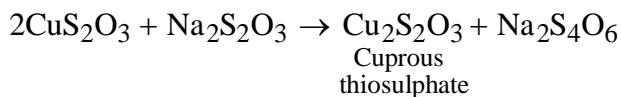
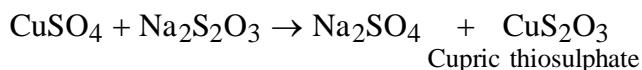
(b) When hypo is in excess : The white ppt. of $\text{Ag}_2\text{S}_2\text{O}_3$ is dissolved in $\text{Na}_2\text{S}_2\text{O}_3$ due to complex formation.



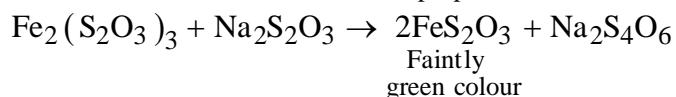
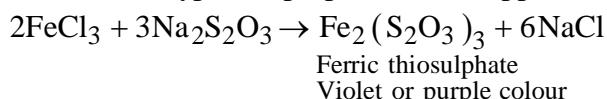
(vi) Reaction with silver halides : Silver halides (AgCl , AgBr and AgI) dissolve in hypo due to complex formation.



(vii) Reaction with copper sulphate : Cuprous thiosulphate is formed which dissolves in excess of hypo to form a complex



(viii) Reaction with ferric chloride : A violet or purple colour is obtained on adding hypo to FeCl_3 solution. In the presence of excess of hypo, the purple colour disappears.



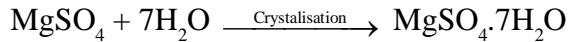
6. Magnesium sulphate (Epsom salt) $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$:

Preparation :

(i) By dissolving magnesite or dolomite in dilute H_2SO_4 .



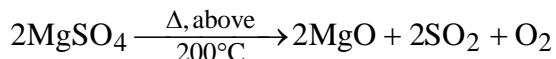
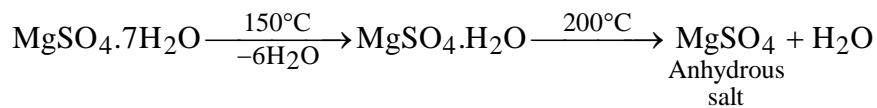
In the second case, white ppt. of CaSO_4 is filtered off and the filtrate on crystallization gives $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$.



Properties :

It is a colourless, crystalline solid, soluble in water with evolution of heat. It loses its water of crystallization on exposure to air, i.e., it shows efflorescence. It has a bitter taste.

Effect of temperature :



With alkali sulphates, it forms double salts of type $\text{K}_2\text{SO}_4 \cdot \text{MgSO}_4 \cdot 6\text{H}_2\text{O}$.

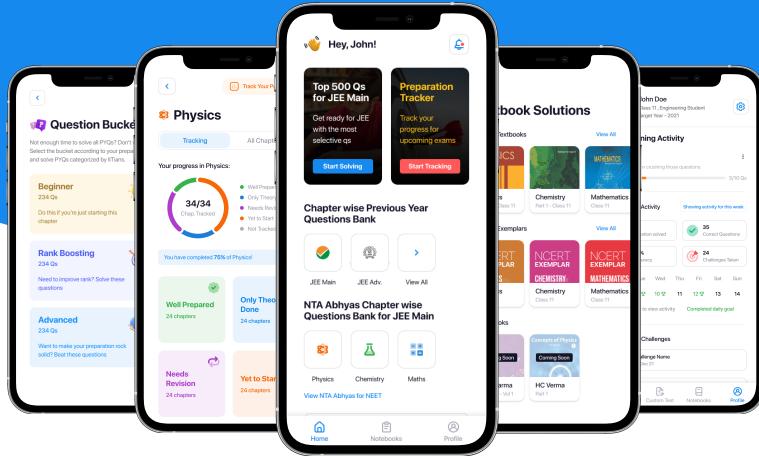
- In $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$, six H_2O molecules are attached to Mg^{+2} while the seventh is attached with SO_4^{2-} .

Uses :

- (i) In medicine as a purgative.
- (ii) In sizing and loading paper, silk and cotton.
- (iii) In dyeing and tanning processes.
- (iv) In ceramic, cement and match industries.
- (v) In the treatment of textile fibres to make them non-inflammable.
- (vi) As stimulant to increase bile secretions.



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