

act

S-Block Elements

① Last e- in s-orbital.

→ s-orbital accommodate 2e-  
two Groups - ① and ②.

→ Group 1 and Group 2

↓  
Alkali Metals

Alkaline Earth Metals

On reaction with H<sub>2</sub>O form strongly alkaline hydroxides

[with exception of Be] oxides and hydroxide are alkaline and found in earth's crust.

→ Abundance :-

(i) Na and K are abundant  
Li, Rb, Cs → lower abundance  
Po radioactive - " longest lived isotope 223 for " (min)

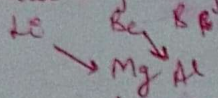
(ii) Ca and Mg rank 5<sup>th</sup> and 8<sup>th</sup> in abundance on earth's crust. Sr, Ba have lower abundance. Be is rare and Ra is rarest comprising only 10-10% of igneous rocks.

→ General EC :-

Alkali [Noble Gas] ns<sup>1</sup>  
Alkaline earth [Noble Gas] ns<sup>2</sup>

→ Diagonal Relationship

first element of Group 1 and Group 2 share some properties with the second element of the following group.



Diagonal relationship is due to similarities in ionic sizes and / or charge/radius ratio of the elements.

→ Na<sup>+</sup>, K<sup>+</sup>, Mg<sup>2+</sup>, Ca<sup>2+</sup> are present in large amounts in biological fluids to function by maintaining ion balance and nerve impulse induction.

② GROUP I (Alkali Metals)

(i) EC → [Ne] ns<sup>1</sup>

loosely bound one e<sup>-</sup> makes them the "Most electropositive metals". They readily form M<sup>+</sup> ion and thus never occur in free state in nature.

(ii) Atomic and ionic sizes increase down the group

$R(M^+) < R(M)$

(iii) Very low ionization enthalpy. Down the group (Li - Cs) decrease.

As increasing size outweighs the increase in nuclear charge.

(iv) Hydration enthalpy

$\Delta H^+ > Na^+ > K^+ > Rb^+ > Cs^+$

side mostly hydrated (e.g. LiCl · 2H<sub>2</sub>O)

(v) Physical Properties

- Silvery white, soft, light
- Density → low due to large size

Increases from Li to Cs

Potassium is lighter than sodium.

→ low MP and BP owing to their weak metallic bonding (due to only single valence e<sup>-</sup>).

→ Imparts colour to their flame. Excites outer e<sup>-</sup> to higher energy level and when the excited e<sup>-</sup> comes back to ground state emitting radiation.

Li	Crimson Red
Na	Yellow
K	Violet
Rb	Red Violet
Cs	Blue

→ They absorb light energy and atom loses e<sup>-</sup>. This property is used in electrodes in photoelectric cell ⇒ Cs and K are used

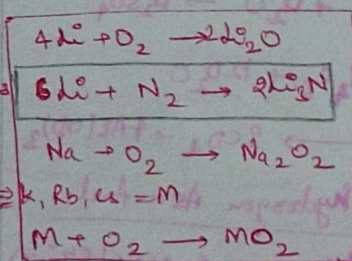
(vi) Chemical Properties

Highly reactive due to large sizes and low IE.

→ Reactivity towards Air

→ Turnish in dry air, forms oxide layer which reacts with moisture forming hydroxides.

⇒ Burns vigorously in air forming oxides.

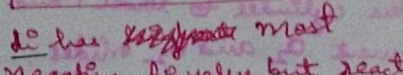


(O<sub>2</sub><sup>-</sup> is stable only in presence of large cations.)

OS of Alkali metals → +1 (low)

Normally kept in kerosene oil

→ Reactivity towards Water

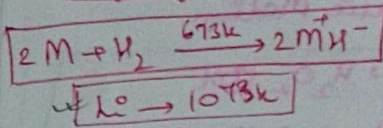


Li has most negative E<sup>o</sup> value but reacts less vigorously with water than sodium (Na) which has least negative E<sup>o</sup> among Group I.

This behaviour is due to its small size, high hydration energy.

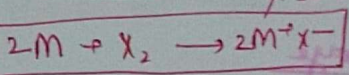
⇒ Also react with protic donors [alcohols, gaseous ammonia, alkynes].

→ Reactivity towards dihydrogen



some high melting solids hydrides are formed.

→ Reactivity towards halogens



However, Lithium halides are covalent in nature.

Li has small size and distorts e<sup>-</sup> cloud of large halide ions. (Polarisation of X<sup>-</sup> takes place).

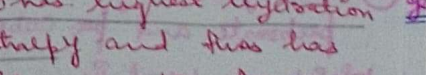
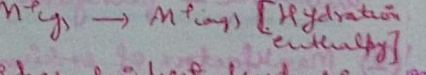
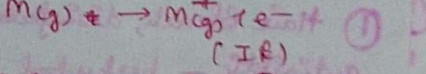
⇒ LI is most covalent.

→ Reactivity towards

→ Reducing Nature

Alkali metals - strong reducing in nature.

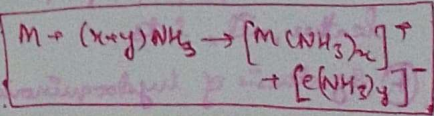
E<sup>o</sup> value determines reducing power.



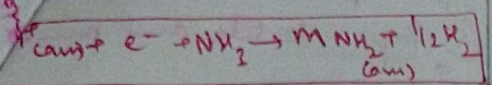
Li has highest hydration enthalpy and has highest negative E<sup>o</sup> value and high reducing power.

→ Solutions in liquid

Ammonia



deep blue in colour  
 ammoniated e-  
 paramagnetic  
 standing slowly liberates  
 to form amide.



In concentrated solutions,  
 blue  $\rightarrow$  Bronze colour and  
 becomes diamagnetic.

(vii) Uses

① Li  $\rightarrow$  Used to make alloys  
 with lead  $\rightarrow$  white metal  
 bearings of motor engine  
 Aluminium  $\rightarrow$  aircraft parts  
 magnesium  $\rightarrow$  armour plates  
 Used in thermonuclear reactions  
 Used to make electrochemical  
 cells.

② Na  $\rightarrow$  Used to make Na/Pb  
 alloy which is  
 needed to make PbEt<sub>4</sub>, PbMe<sub>4</sub>  
 which were used as  
 anti-knock additives in  
 petrol (but not now - lead free  
 petrol),  
 Liquid Na metal used as  
 coolant in fast breeder  
 nuclear reactor.

③ K  $\rightarrow$  vital role in biological  
 system.  
 KCl used as fertilizer  
 KOH used in manufacture of  
 soft soap and used as an  
 excellent absorbent of CO<sub>2</sub>.

④ Cs  $\rightarrow$  Used in photoelectric  
 cells.

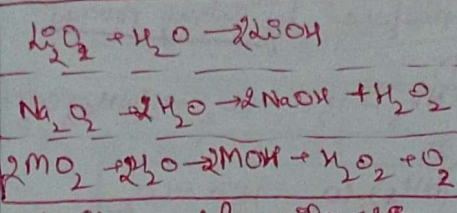
③ General characteristics of  
 Compounds of Alkali Metals

① Oxides and Hydroxides

On combustion,  
 Li forms oxide (and some  
 peroxide), Na forms peroxide  
 and some superoxide) and  
 K, Rb, Cs forms superoxide.

Increasing stability of  
 peroxide or superoxide is  
 due to stabilization of  
 large ~~oxo~~ anions by large  
 cations through their lattice  
 energy.

Hydrolysed :-



Oxides and peroxides are  
 colourless but superoxides  
 are yellow <sup>orange</sup> in colour.

Superoxides are paramagnetic  
 $\rightarrow$  Na<sub>2</sub>O<sub>2</sub> used widely as an  
 oxidising agent.

Hydroxides formed are  
 white crystalline solids.  
 $\rightarrow$  evolve large amount of  
 heat on dissolving water  
 $\rightarrow$  strongest bases.

Q Why KO<sub>2</sub> is paramagnetic?  
 = Because it has one  
 unpaired e- in n\*2p orbital.

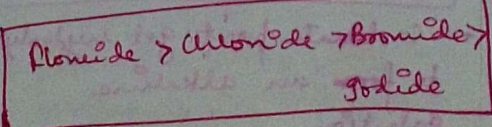
(ii) Halides

$\rightarrow$  High melting, crystalline,  
 colourless solids.  
 $\rightarrow$  MX + ~~oxide~~ hydroxide/  
 carbonate.

$\rightarrow$  For a given metal,  
 $\Delta H^\circ \equiv$  less negative  
 from fluoride to iodide.

$\rightarrow$  For fluoride,  
 $\Delta H^\circ$  becomes less negative  
 down the group.

$\rightarrow$  For chlorides, bromides  
 and iodides,  
 $\Delta H^\circ$  become less negative  
 up the group.  
 $\rightarrow$  MP and BP



$\rightarrow$  All these halides are soluble  
 soluble in water.

$\rightarrow$  Low solubility of (in H<sub>2</sub>O)  
 LiF  $\Rightarrow$  High lattice ~~energy~~

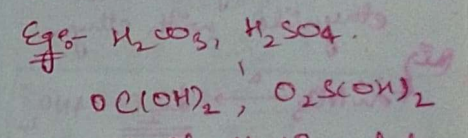
CsI  $\Rightarrow$  Smaller hydration  
 enthalpy

$\rightarrow$  Lithium halides are  
 soluble in ~~ether~~ ethanol,  
 acetone and ethyl acetate.

KCl is also soluble in  
pyridine.

(iii) Salts of Oxo-Acids

$\Rightarrow$  Oxo-Acids = those in  
 which acidic proton is  
 attached on hydroxyl group  
 with an oxo-group attached  
 to same atom.



$\rightarrow$  Salts of alkali metals  
 with oxo-acids are  
 soluble in water and  
 thermally stable.

$\rightarrow$  Down the group,  
 stability of carbonates  
 and hydrogencarbonates  
 increases.  
 [electropositive character]

$\rightarrow$  Lithium carbonate is  
 not stable to heat  
 and forms much stable  
 $\text{Li}_2\text{O}$ ,  $\text{CO}_2$

Li<sup>+</sup> (small size) polarises  
 large CO<sub>3</sub><sup>2-</sup> ion

\* Lithium hydrogencarbonate  
 does not exist as solid.

④ Anomalous properties of Li

- ① Small size
- ② High polarising power  
 (i.e. charge/radius ratio)  
 $\rightarrow$  lithium compounds  
 have increased covalent  
 character.