→ used.

— × — × — ×

S-Block Elements

(1) Last e^- in S-orbital.

→ S-orbital accomodate $2e^-$, two Groups — 1 and 2.

→ Group 1 and Group 2

Alkali Metals	Alkaline Earth Metals
On reaction with H_2O 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. Fr radioactive — " Longest lived isotope 223 for " ($t_{1/2}$ = 21 min)

(ii) Ca and Mg rank 5^{th} and 6^{th} 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^1

Alkaline Earth [Noble Gas] ns^2

→ Diagonal Relationship

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

Li Be B
↘ Mg Al

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

→ Na^+, K^+, Mg^{2+}, Ca^{2+} 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 → [NG] ns^1

loosely bound one e^- makes them the "Most Electropositive metals". They readily form M^+ ion and thus never occur in free state in nature.

(ii) Atomic and ionic sizes increase down the group

$\boxed{R(M^+) < R_M}$

(iii) Very low ionization enthalpy. Down the group (Li – Cs) decrease.
As increasing size outweighs the increase in nuclear charge.

(iv) Hydration Enthalpy

$\boxed{Li^+ > Na^+ > K^+ > Rb^+ > Cs^+}$

Salts mostly hydrated [Ex, $LiCl \cdot 2H_2O$]

(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^-)

→ Imparts colour to their flame
[Excites outer e^- to higher energy level and when the excited e^- 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^-.
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:–

1. Tarnish in dry air, forms oxide layer which reacts with moisture forming hydroxides.

2. Burns vigorously in air forming oxides.

$4Li + O_2 \rightarrow 2Li_2O$

$\boxed{6Li + N_2 \rightarrow 2Li_3N}$

$Na + O_2 \rightarrow Na_2O_2$

K, Rb, Cs = M

$M + O_2 \rightarrow MO_2$

(O_2^- is stable only in presence of large cations.)

OS of Alkali metals → +1 (here)
Normally kept in kerosene oil

→ Reactivity towards Water:–

$\boxed{M + H_2O \rightarrow 2M^+ + 2OH^- + H_2}$

1. Li has ~~highest~~ most negative E° value but reacts less vigorously with water than sodium (Na) which has least negative E° among Group I.

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

2. Also react with proton donors [alcohols, gaseous ammonia, alkynes].

→ Reactivity towards dihydrogen:–

$\boxed{2M + H_2 \xrightarrow{673K} 2M^+H^-}$

$\boxed{Li \rightarrow 1073K}$

Ionic | High melting solids hydrides are formed.

→ Reactivity ~~for~~ towards halogens:–

$\boxed{2M + X_2 \rightarrow 2M^+X^-}$ ↗ Ionic

However,

Lithium halides are covalent in nature

↓

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

⇒ LiI is most covalent

~~Reactivity towards~~

→ Reducing Nature:–

Alkali metals – Strong reducing in nature.

E° value determines reducing power.

$M(s) \rightarrow M(g)$ (Sublimation)

$M(g) \rightarrow M^+(g) + e^-$ (IE)

$M^+(g) \rightarrow M^+(aq)$ [Hydration enthalpy]

Li has highest hydration enthalpy and thus has highest negative E° value and high reducing power.

→ Solutions in liquid Ammonia:–

$\boxed{M + (x+y)NH_3 \rightarrow [M(NH_3)_x]^+ + [e(NH_3)_y]^-}$

... deep blue in colour
... ammoniated e^-
... paramagnetic
... standing slowly liberates H_2 to form amide

$M^+_{(am)} + e^- + NH_3 \to MNH_{2(am)} + \frac{1}{2}H_2$

* In concentrated solutions, blue → Bronze colour and becomes diamagnetic.

(vii) Uses

(1) Li → Used to make alloys with lead → white metal, bearings of motor engine
Aluminium → aircraft parts
Magnesium → armour plates
Used in thermonuclear reactions
Used to make electrochemical cells.

(2) Na → Used to make Na/Pb alloy which is needed to make $PbEt_4$, $PbMe_4$ 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 → vital role in biological system.
KCl used as fertilizer
KOH used in manufacture of soft soap and used as an excellent absorbent of CO_2.

(4) Cs → Used in photoelectric cells.

(3) General characteristics of Compounds of Alkali Metals:-

(i) 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 ~~cations~~ anions by large cations through their lattice energy.

Hydrolysed :-

$Li_2O + H_2O \to 2LiOH$

$Na_2O_2 + 2H_2O \to 2NaOH + H_2O_2$

$2MO_2 + 2H_2O \to 2MOH + H_2O_2 + O_2$

→ Oxides and Peroxides are colourless but superoxides are yellow ~~or~~ orange in colour.
→ Superoxides are paramagnetic
→ Na_2O_2 used widely as an oxidising agent.
→ Hydroxides formed are white crystalline ~~solids~~.
→ Evolve large amount of heat on dissolving water
→ strongest bases.

Q Why KO_2 is paramagnetic?
= Because it has one unpaired e^- in π^*2p orbital.

(ii) Halides
→ High melting, crystalline, colourless solids.
→ HX + oxide/hydroxide/carbonate.
→ For a given metal,
ΔH° = less negative from floride to iodide.
→ for floride,
ΔH° becomes less negative down the group.
→ for chlorides, Bromides and iodides,
ΔH° become less negative up the group.
→ MP and BP

Floride > Chloride > Bromide > Iodide

→ All these halides are soluble in water
→ Low solubility of (in H_2O)
LiF ⇒ high lattice enthalpy
CsI ⇒ Smaller hydration enthalpy
→ Lithium halides are soluble in ~~alcohol~~ ethanol, acetone and ethyl acetate. LiCl is also soluble in pyridine.

(iii) Salts of Oxo-Acids
⇒ Oxo Acids :- Those in which acidic proton is on hydroxyl group with an oxo - group attached to same atom.
Eg:- H_2CO_3, H_2SO_4.
$OC(OH)_2$, $O_2S(OH)_2$
→ Salts of alkali metals with oxo - acids are soluble in water and thermally stable.
→ Down the group, Stability of carbonates and hydrogencarbonates increases.
[electropositive character ↑]
→ Lithium carbonate is not stable to heat and forms much stable Li_2O, CO_2
Li^+ (small size) polarises large CO_3^{2-} 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), → Lithium compounds have increased Covalent character.