8. Elements of Groups 1 and Group 2

We have seen in chapter 7 that the element of group 1 and group 2 belong to the s-block of the modern periodic table.

Can you recall?

1. Which is the first element in the periodic table ?
2. What are isotopes ?
3. Write the formulae of the compounds of hydrogen formed with sodium and chlorine.

Hydrogen is the first element in the periodic table. Hydrogen appears at the top of group 1 of the alkali metals. But it differs from alkali metals in a number of respects and, therefore, is studied separately.

8.1 Hydrogen : Hydrogen has the simplest atomic structure of all the elements. A hydrogen atom consists of a nucleus of charge +1 and one extranuclear electron. Hydrogen has a little tendency to lose this electron however can pair with the other electron easily forming a covalent bond. It exists in diatomic form as H_2 molecule; therefore it is often called **dihydrogen**.

8.1.1 Occurrence : In the free state hydrogen exists as dihydrogen gas. Hydrogen is most abundant element in the universe; it makes 70 % of the total mass of the universe. Hydrogen is also the principal element in the solar system. On the earth, hydrogen is the tenth most abundant element on mass basis and the third most abundant element on atom basis.

8.1.2 Position of hydrogen in the periodic table : Position of hydrogen in the periodic table has been a subject of discussion.

Electronic configuration of hydrogen is $1s^1$ which is similar to ns^1 which is the outer electronic configuration of alkali metals belonging to the group 1. But $1s^1$ also resembles the outer electronic configuration of group 17 elements, which is ns^2np^5. By adding one electron to H we get electronic configuration of the inert gas He which is $1s^2$ and by adding one electron to ns^2np^5 we get ns^2np^6 which is the outer electronic configuration of the remaining inert gases. Some chemical properties of hydrogen are similar to those of alkali metals while some resemble halogens. The uniqueness of hydrogen is that $H^{\oplus}$ formed by loss of the electron from hydrogen atom does not exist freely. It is always associated with other molecules. For example:

$$H^{\oplus} + H_2O \longrightarrow H_3O^{\oplus}$$

This is because $H^{\oplus}$ is nothing else but a proton. Hydrogen is, therefore, placed separately above the group 1. It may be noted, here, that metastable metallic hydrogen was discovered at Harvard university, USA, in January 2017.

8.1.3 Isotopes of Hydrogen : If different atoms of the same element have different mass numbers they are called isotopes of each other. (Refer to chapter 4). Hydrogen has three isotopes with mass numbers 1, 2 and 3. They all contain one proton and one electron but different number of neutrons in the nucleus. (see figure 8.1 and table 8.1)

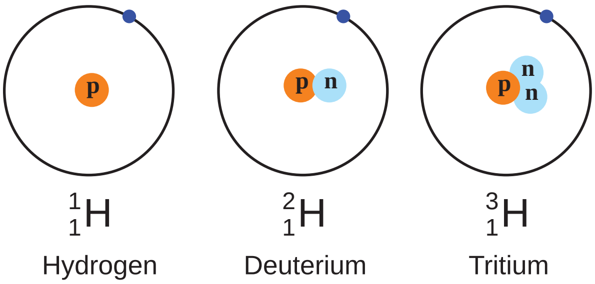


Figure 8.1 Isotopes of hydrogen

Can you tell?

In which group should hydrogen be placed ? In group 1 or group 17 ? Why ?

Table 8.1 : Characteristics of isotopes of hydrogen

Name of the Isotope	Symbol	Atomic number *Z*	Atomic mass number *A*	Neutron number *N*	Abundance	Stability
Hydrogen	H or ${}^{1}_{1}H$ or H-1	1	1	0	99.98 %	Stable
Deuterium	D or ${}^{2}_{1}H$ or H-2	1	2	1	0.015 %	Stable
Tritium	T or ${}^{3}_{1}H$ or H-3	1	3	2	Trace	Ratio active

Do you know ?

Tritium ${}^{3}_{1}H$ is a radioactive nuclide with half life period 12.4 years and emits low energy $\beta^{\ominus}$ particles.

8.1.4 Preparation of dihydrogen

Hydrogen can be prepared using many methods.

A. Laboratory methods

i. Dihydrogen is prepared in laboratory by the action of dilute hydrochloric acid on zinc granules.

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

ii. Dihydrogen can be prepared by the action of aqueous solution of sodium hydroxide on zinc.

$$Zn(s) + 2NaOH(aq) \rightarrow Na_2ZnO_2(aq) + H_2(g)$$

B. Industrial methods

i. By electrolysis of pure water: Pure water is a poor conductor of electricity. Therefore a dilute aqueous solution of acid or alkali is used to prepare dihydrogen by electrolysis. For example, electrolysis of dilute aqueous solution of sulphuric acid yields two volumes of hydrogen at cathode and one volume of oxygen at anode.

$$2H_2O\ (l) \xrightarrow[\text{trace of acid or akali}]{\text{electrolysis}} 2H_2\uparrow + O_2\uparrow$$

Pure dihydrogen (> 99.5% purity) gas is obtained by electrolysis of warm solution of barium hydroxide between nickel electrodes.

ii. From carbon or hydrocarbon: Three stages are involved in this industrial process of preparation of dihydrogen.

Stage 1 : Reaction of steam on hydrocarbon or coke (C) at 1270 K temperature in presence of nickel catalyst, gives water-gas, which is a mixture of carbon monoxide and hydrogen.

$$CH_4(g) + H_2O(g) \xrightarrow[\text{Ni}]{1270\ \text{K}} \underbrace{CO(g) + 3H_2(g)}_{\text{water-gas}}$$

As water-gas is used for synthesis of CH_3OH and a number of hydrocarbons, it is also called **'syngas'**. Production of syngas is also the first stage of gasification of coal.

$$C(s) + H_2O\ (g) \xrightarrow{1270\ \text{K}} CO(g) + H_2(g).$$

Sawdust, scrapwood, etc. can also be used in place of carbon.

Stage 2 : Water-gas shift reaction: The carbon monoxide in the watergas is transformed into carbondioxide by reacting with steam in presence of iron chromate as catalyst. This is called water-gas shift reaction.

$$CO(g) + H_2O(g) \xrightarrow[\text{iron chromate catalyst}]{673\ \text{K}} CO_2(g) + H_2(g)$$

Stage 3 : Carbon dioxide is removed by scrubbing with sodium arsenite solution.

Today major industrial production of dihydrogen is from petrochemicals (~ 77 %), about 18 % from coal, about 4 % by electrolytic methods and about 1 % by other methods.

8.1.5 Properties of dihydrogen

A. Physical properties : Dihydrogen is colourless, tasteless and odourless gas. It burns with a pale blue flame. It is a nonpolar water insoluble gas, lighter than air.

B. Chemical properties :

i. Reaction with metals : Dihydrogen combines with all the reactive metals such as alkali metals, calcium, strontium and barium at high tempereture, to form metal hydrides. For example :

$$2Na(s) + H_2(g) \xrightarrow{\Delta} 2NaH(s)$$

(In this respect dihydrogen is similar to halogens which also react with metals and form metal halides.)

Just think

In the above chemical reaction which element does undergo oxidation and which does undergo reduction ?

Do you know ?

The bond dissociation energy of H-H bond is high, which is 436 kJ mol^{-1}. Therefore reactions of dihydrogen take place at high temperature and /or in the epresence of catalyst.

ii. Reaction with dioxygen: Dihydrogen reacts with dioxygen in the presence of catalyst or by heating to form water. This reaction is highly exothermic.

$$2H_2(g) + O_2(g) \xrightarrow[\text{or heating}]{\text{catalyst}} 2H_2O(l) \; ; \quad \Delta H = -235 \text{ kJ mol}^{-1}$$

iii. Reaction with halogens: Dihydrogen inflames with fluorine even at -250^0 C in dark, whereas it requires catalyst to react with iodine. The vigour of reaction of dihydrogen decreases with increasing atomic number of halogen.

$$H_2\,(g) + X_2(g) \rightarrow 2HX(g)$$

(In this respect dihydrogen resembles alkali metals which also react with halogens to form halides.)

iv. Reducing nature of dihydrogen :

a. Dihydrogen reduces oxides and ions of metals those are less reactive than iron, to the corresponding metals at moderate temperature. For example :

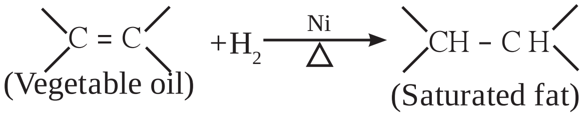
$$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l)$$

$$Fe_3O_4\,(s) + 4\,H_2(g) \longrightarrow 3Fe(s) + 4H_2O(l)$$

$$Pd^{2\oplus}(aq) + H_2\,(g) \longrightarrow Pd(s) + 2H^{\oplus}(aq)$$

b. Hydrogenation of unsaturated organic compound

The hydrogenation of unsaturated organic compounds such as oil using nickel caralyst gives saturated organic compounds such as solid fat (vanaspati Ghee).



7.1.6 Uses of dihydrogen

i. Largest use of dihydrogen is in production of ammonia.

ii. Dihydrogen is used in formation of vanaspati ghee by catalytic hydrogenation of oils.

iii. Liquid dihydrogen is used as a rocket fuel.

iv. Dihydrogen is used in preparation of important organic compounds like methanol in bulk quantity.

$$2H_2(g) + CO(g) \xrightarrow[\text{Cobalt catalyst}]{} CH_3OH(l)$$

v. Dihydrogen is used for preparation of hydrogen chloride (HCl) and metal hydrides.

Problem 8.1 : Justify the placement of hydrogen in the group of alkali metals with the help of reaction with halogens.

Solution :

Hydrogen on reaction with halogen (X_2) gives compounds with general formula HX. For example : $H_2 + Cl_2 \longrightarrow 2HCl$ similarly alkali metals (M) on reaction with halogens (X_2) give compounds with general formula MX.

For examples : $2Na + Cl_2 \longrightarrow 2NaCl$

Thus, H_2 and alkali metal are monovalent elements more electropositive than halogens. This similarily justifies the place of hydrogen in the group 1.

Can you recall?

1. What is the name of the family of reactive metals having valancy one?
2. What is the name of the family of reactive metals having valancy two?

8.2 Alkali metals and alkaline earth metals

8.2.1 Introduction : The elements of the groups 1 and 2 are placed on the left in the periodic table. Here the last electron enters into '*ns*' subshell. Thus they belong to the s-block of the periodic table. **Group 1** of the periodic table consists of the elements: **hydrogen, lithium, sodium, potassium, rubidium, caesium and francium**. These elements **except hydrogen** are collectively called **alkali metals**. We have already looked at hydrogen in the section 8.1. Two elements of Group 1, namely, sodium and potassium are the sixth and seventh most abundant elements in earth crust but francium does not occur appreciably in nature because it is radioactive and has short half-life period.

Group 2 of the periodic table consists of elements : **beryllium, magnesium, calcium, strontium, barium and radium**. These elements are collectively called **alkaline earth metals** because they occur as minerals in rocks. The elements magnesium and calcium are found abundantly in earth crust but radium is not easy to find. Radium is one of the first two radioactive elements discovered by Madame Curie.

8.2.2 Electronic configuration of elements of group 1 and group 2

The general outer electronic configuration of the group 1 elements is ns^1 and that of the group 2 elements is ns^2. The loosely held s-electrons in valence shell of these elements can be easily removed to form metal ions. Hence these elements are never found in free state in nature. Tables 8.2 and 8.3 show the electronic configurations of the elements of group 1 and group 2 elements, respectively.

Table 8.2 : Electronic configuration of group 1 elements

Name	Symbol	Atomic number	Condensed electronic configuration	Electronic Configuration.
Hydrogen	H	1	$1s^1$	$1s^1$
Lithium	Li	3	$[He]\ 2s^1$	$1s^2\ 2s^1$
Sodium	Na	11	$[Ne]\ 3s^1$	$1s^2\ 2s^2\ 2p^6\ 3s^1$
Potassium	K	19	$[Ar]\ 4s^1$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
Rubidium	Rb	37	$[Kr]\ 5s^1$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^1$
Caesium	Cs	55	$[Xe]\ 6s^1$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^1$
Francium	Fr	87	$[Rn]7s^1$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6\ 7s^1$

Table 8.3 : Electronic configuration of group 2 elements

Name	Symbol	Atomic number	Condense electronic configuration	Electronic Configuration.
Beryllium	Be	4	$[He]\ 2s^2$	$1s^2\ 2s^2$
Magnesium	Mg	12	$[Ne]\ 3s^2$	$1s^2\ 2s^2\ 2p^6\ 3s^2$
Calcium	Ca	20	$[Ar]\ 4s^2$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$
Strontium	St	38	$[Kr]\ 5s^2$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2$
Barium	Ba	56	$[Xe]\ 6s^2$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2$
Radium	Ra	88	$[Rn]\ 7s^2$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6\ 7s^2$

8.2.3 Trends in atomic and physical properties of elements of group 1 and group 2

All the **alkali metals** are silvery white and soft. Due to their large atomic size these elements have low density. They are the most electropositive elements. The **alkaline earth** are **metals** also in general silvery white lustrous and soft, but harder than the alkali metals. They are also strongly electropositive in nature. But comparatively less electropositive than the alkali metals. Some atomic and physical properties of the alkali metals and the alkaline earth metals are listed in the tables 8.4 and 8.5 respectively.

Table 8.4 : Physical properties of group 1 elements (except hydrogen)

Symbol	Atomic radius (pm)	Ionic radius (pm)	Density (g/cm^{-3})	Ionization enthalpy (kJ mol^{-1})	Electro-negativity	Melting point (K)	Abundance in the lithosphere	Standard reduction potential E^o (V)
Li	152	76	0.54	520	1.0	454	18 ppm	-3.04
Na	186	102	0.97	496	0.9	371	2.27 %	-2.714
K	227	138	0.86	419	0.8	336	1.84 %	-2.925
Rb	248	152	1.53	403	0.8	312	78.12 ppm	-2.930
Cs	265	167	1.90	376	0.7	302	2.6 ppm	-2.927
Fr	-	(180)	-	~375	-	-	-10^{-18} ppm	-

Table 8.5 : Physical properties of group 2 elements

Symbol	Atomic radius (pm)	Ionic radius (pm)	Density (g/cm^{-3})	Ionization enthalpy (kJ mol^{-1})		Electro-negativity	Melting point (K)	Abundance in the lithosphere	Standard reduction potntial E^0 (V)
				1st	2nd				
Be	111	31	1.84	899	1757	1.5	1560	2 ppm	-1.97
Mg	160	72	1.74	737	1450	1.2	924	2.76 %	-2.36
Ca	197	100	1.55	590	1145	1.0	1124	4.6 %	-2.84
Sr	215	118	2.63	549	1064	1.0	1062	384 ppm	-2.89
Ba	222	135	3.59	503	965	0.9	1002	390 ppm	-2.92
Ra	-	148	(5.5)	509	979	-	973	10^{-6} ppm	-2.92

Unipositive ions of all the elements of group 1 have inert gas configuration. Thus they have no unpaired electron and their compounds are diamagnetic and colourless. The divalent ions of group 2 elements also have inert gas configuration with no unpaired electron, and therefore their compounds are also diamagnetic and colourless. Lithium and beryllium differ from the rest of the elements of the groups 1 and 2, respectively because of their extremely small size and comparatively high electronegativity.

The physical properties of group 1 and group 2 elements show reasonable regularity in the periodic trends. Thus the atomic and ionic radii and densities increase down both the groups. Ionization enthalpies and electronegativities decrease down both the groups. The elements of both these groups, in general, have high negative values of standard reduction potentials.

8.2.4 Chemical properties of elements of group 1 and group 2

The alkali metals and alkaline earth metals are very reactive in nature. As a result of this they are always found in combined state. Their reactivity is due to their low ionization ehthalpy values in general. The reactivity of these metals increases with increasing atomic radius and corresponding lowering of lonization enthalpy down the groups 1 and 2 can be noticed.

Problem 8.2:
Sodium forms ionic compounds having formulae NaCl, NaH and Na_2CO_3. Explain

Solution: Let us rewrite the formulae of the compounds of sodium showing charges on the concerned cation and basic anion.

$Na^{\oplus}Cl^{\ominus}$, $Na^{\oplus}H^{\ominus}$, $2Na^{\oplus}\ CO_3^{2\ominus}$

It is seen that in all these compounds Na carries one positive charge. The $Na^{\oplus}$ is formed from Na atom by losing one electron. $Na \rightarrow Na^{\oplus} + e^{\ominus}$

This happens because the electronic configuration of Na is $[Ne]3s^1$. There is only one electron in the valence shell of Na. It can easily be lost as the ionization enthalpy is low. And $Na^{\oplus}$ ion so formed is stable as it has stable electronic configuration of the inert gas Ne.

Problem 8.3 : Explain the observed values 496 kJ/mol and 737 kJ/mol of the first ionization enthalpies of Na and Mg, respectively.

Solution : The electronic configuration of Na is $[Ne]3s^1$ and that of Mg is $[Ne]3s^2$. During the first iouization only one electron is removed from a neutral atom.

$$Na \xrightarrow{1^{st}\ \text{ionisation}} Na^{\oplus} + e^{\ominus}$$
$$Mg \xrightarrow{1^{st}\ \text{ionisation}} Mg^{\oplus} + e^{\ominus}$$

The resulting Na is isoelectronic with Ne, and therefore, is stable. Thats why Na has low value of 1st ionization enthalpy. However, to form the unipositive $Mg^{\oplus}$ ion energy is required to unpair these electrons in the valency shell and also remove one electron to form the ion having electronic configuration $[Ne]\ 3s^1$. It is not as stable as Na since it does not correspond to any inert gas. Therefore the first ionization enthalpy of Mg is higher than that of Na.

Problem 8.4:
What is the oxidation state of Na in Na_2O_2?

Solution: The peroxide species is represented as O_2^{2-}. Any compound is electrically neutral. Therefore oxidation state of each Na is (+2/2 = +1) in Na_2O_2.

Problem 8.5 :
The atomic radii of Na, K and Mg are 186, 227 and 160 pm, respectively. Explain the differences.

Solution : Na and K both belong to the group 1. K has larger valence shell than Na. Therefore the atomic radius of K is larger than that of Na. Na and Mg belong to the same period. Therefore both have the same valence shell. But the nuclear charge of Mg is larger than that of Na. Therefore, the valence electrons of Mg are held more tightly and its atomic radius is smaller than that of Na.

The elements of group 1 and group 2 both being s-block elements, show similarily in their chemical properties. The differences are due to variation in the atomic radii, ionization enthalpies and valencies.

i. Reaction with oxygen/air

Group 1 - All the elements of group 1 rapidly lose their luster in air due to formation of a layer of oxide, on peroxide and in some cases superoxide by reaction with oxygen in air.

$$2Li + O_2 \longrightarrow 2LiO$$
(Lithium oxide)

$$2Na + O_2 \longrightarrow Na_2O_2$$
(Sodium peroxide)

$$K + O_2 \longrightarrow KO_2$$
(Potassium superoxide)

Do you know ?

- The reaction of Na and K with oxygen is highly exothermic and these metals catch fire when exposed to air.

- Potassium superoxide has ability to absorb carbon dioxide and give out oxygen at the same time:

  $$4KO_2 + 2CO_2 \rightarrow 2K_2CO_3 + 3O_2\uparrow$$

- This property of KO_2 has been made use of in breathing equipment used for mountaineers and in submarines and space.

The oxides of group 1 metals are strongly basic in nature. They dissolve in water forming aqueous solutions of strong alkali. For example

$$2LiO(s) + H_2O(l) \rightarrow 2LiOH(aq)$$

Group 2 : These metals are protected from air oxidation by an oxide film formed on their surface.

All the elements of group 2 burn when ignited in air forming MO type oxides. The product is a mixture of oxide and nitride.

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

$$3Mg + N_2 \rightarrow Mg_3N_2$$

Further heating of the oxide in air results in formation of peroxide.

ii. Reaction with water

Group 1 : Lithium, sodium and potassium all float on water due to hydrogen bubbles released on reaction with water. Lithium reacts slowly but sodium and potassium react vigrously with water. Due to highly exothermic reaction sodium and potassium catch fire when put in water.

$$2Na + 2H_2O \rightarrow 2Na\,OH + H_2\uparrow$$

Group 2 : The elements of group 2 react with water to form metal hydroxide and hydrogen. Beryllium does not react with water. Magnesium decomposes hot water, other elements react with cold water forming metal hydroxide $M(OH)_2$ and hydrogen gas.

$$Ca + 2H_2O \rightarrow Ca(OH)_2\downarrow + H_2\uparrow$$

iii. Reaction with Hydrogen

Group 1 : Alkali metals react with hydrogen at high tempereture to form the corresponding metal hydrides.

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

Group 2 : All the metals of group 2, except beryllium, when heated with hydrogen form MH_2 type hydrides.

$$M + H_2 \xrightarrow{\Delta} MH_2$$

Problem 8.6 : NaCl is an ionic compound but LiCl has some covalent character, explain.

Solution: $Li^{\oplus}$ ion has very small size, therefore the charge density on $Li^{\oplus}$ is high. Therefore it has high tendency to distort the electron cloud around the negatively charged large chloride ion. This results in partial covalent character of the LiCl bond. $Na^{\oplus}$ ion cannot distort the electron cloud of $Cl^{\ominus}$ due to the bigger size of $Na^{\oplus}$ compared to $Li^{\oplus}$.

iv. Reaction with Halogens

Group 1 : All the alkali metals react vigorously with halogens to produce their ionic halide salts.

$$2M + X_2 \longrightarrow 2M^{+}X^{-}$$

Group 2: All the alkaline earth metals combine with halogens at high temperature to form halides.

$$M + X_2 \longrightarrow MX_2$$

v. Reducing nature

Group 1 : The reducing power of an element is measured in terms of standard electrode potential (E^0) corresponding to the transformation $M^{\oplus}(aq) + e^{\ominus} \rightarrow M(s)$. All the alkali metals have high negative values of E^0 indicative of their strong reducing nature, lithium is the most powerful and sodium is the least powerful in the group. (see Table 8.4)

Group 2 : All the alkaline earth metals have high negative values of standard reduction potential (E^0). (see Table 8.5), and are strong reducing agents. However their reducing power is less than those of alkali metals.

vi Solution in liquid ammonia

Group 1 : The alkali metals are soluble in liquid ammonia giving deep blue coloured solutions which show electrical conductivity.

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

The ammoniated electron is responsible for the deep blue colour of these solutions. These solutions are paramagnetic and on standing liberate hydrogen slowly, resulting in formation of the metal amide. The blue colour changes to bronze and the solution becomes diamagnetic.

$$M^{\oplus}(am) + e^{\ominus} (am) + NH_3 (l) \rightarrow MNH_2(am) + H_2(g)^{\ominus}$$

(Here (am) denotes solution in ammonia.)

Group 2 : Similar to alkali metals the alkaline earth metals are also soluble in liquid ammonia which give deep blue black coloured solutions.

$$M + (x + 2y) NH_3 \rightarrow [M(NH_3)_x]^{2\oplus} + 2[e(NH_3)_y]^{\ominus}$$

8.2.5 Diagonal Relationship : It is expected that elements belonging to the same group exhibit similarity and gradation in their properties. The first alkali metal lithium and the first alkaline earth metal beryllium do not fulfil this expectation. Thus, lithium shows many differences when compared with the remaining alkali metals and resembles with magnesium, the second alkaline earth metal. Likewise beryllium shows many differences with remaining alkaline earth metals and shows similarity with aluminium, the second element of the next main group (group 13). The relative placement of these elements with similar properties in the periodic table appears to be across a diagonal (see. Table 8.6) and is called diagonal relationship.

Table 8.6 : Diagonal relationship

Main Group / Period	1	2	13
2	Li	Be	B
3	Na	Mg	Al

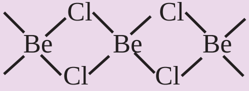
The table 8.7 shows some properties of lithium and magnesium which elucidate their diagonal relationship.

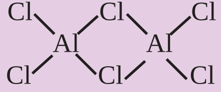
Table 8.7 : Ressemblence between Li and Mg

Criterion / Element	Products of reaction with air $M \xrightarrow{air} ?$	Products of thermal decomposition of carbonate $M_2CO_3/MCO_3 \xrightarrow{\Delta} ?$	Property of chloride	Formula of crystalline chloride
Li	$Li_2O + Li_3N$	$Li_2O + CO_2$	deliquescent	$LiCl.2H_2O$
Mg	$MgO + Mg_3N_2$	$MgO + CO_2$	deliquescent	$MgCl_2.8H_2O$
Group 1 (except Li)	$M_2O/M_2O_2/MO$	No reaction	Not deliquescent	MCl
Group 2	$MO + M_3N_2$	$MO + CO_2$	deliquescent	$MCl_2.xH_2O$

In table 8.8 some properties of Be and Al are shown which indicate the diagonal relationship

Table 8.8 : Resemblence between Be and Al

Criterion / Element	Properties of chloride			Properties of oxide Acidic/Basic/Amphoteric
	Nature of bonding	Whether Lewis acid	Solubility in organic solvent	
Be	**Covalent** chain structure with Cl bridges  	$BeCl_2$ is strong **Lewis acid**	**Soluble**	**Amphoteric** $BeO+2HCl \longrightarrow BeCl_2+H_2$ $BeO+2NaOH \rightarrow Na_2BeO_2+H_2O$

Al	**Covalent** dimer with Cl bridges 	$AlCl_3$ is strong **Lewis acid**	**Soluble**	**Amphoteric** $Al_2O_3 + 6HCl \rightarrow 2AlCl_3+3H_2O$ $Al_2O_3+2NaOH \rightarrow 2NaAlO_2+H_2O$
Group 2	**Ionic**	**Not Lewis acid**	**Insoluble**	**Basic** $MO+HCl \rightarrow MCl_2 + H_2$ $MO + NaOH \rightarrow$ No reaction
Group 13	**Covalent**	**Lewis acid**	**Soluble**	*Amphoteric*

The diagonal relationship between the elemental pairs belonging to different groups and periods is due to the similarity in some of their atomic properties. Thus atomic and ionic radii of Li and Mg are very similar. (see the table 8.4). In the case of Be and Al the charge to radius ratio of their ions is very similar. (Be : $\frac{2}{31}$ and Al : $\frac{3}{53.55}$)

8.2.6 Uses of elements of group 1 and group 2

Group 1

i. Lithium metal is used in long-life batteries used in digital watches, calculators and computers.

ii. Liquid sodium has been used for heat transfer in nuclear power station.

iii. Potassium chloride is used as a fertilzer.

iv. Potassium is used in manufacturing potassium superoxide (KO_2) for oxygen generation. It is good absorbent of carbon dioxide.

v. Caesium is used in photoelectric cells.

Group 2

i. Beryllium is used as a moderator in nuclear reactors.

ii. Alloy of magnesium and aluminium is widely used as structural material and in aircrafts.

iii. Calcium ions are important ingredient in biological system, essential for healthy growth of bones and teeth.

iv. Barium sulphate is used in medicine as barium meal for intestinal x-ray.

v. Radium is used in radiotherapy for cancer treatment.

8.2.7. Biological importance of elements of group 1 and group 2

Group 1

i. Sodium ion is present as the largest supply in all extracellular fluids. These fluids provide medium for transportting nutrients to the cells.

ii. The concentration of sodium ion in extracellular fluids regulates the flow of water across the membrane.

iii. Sodium ions participate in the transmission of nerve signals.

iv. Potassium ions are the most abundant ions within cells. These are required for maximum efficiency in the synthesis of proteins and also in oxidation of glucose.

Group 2

i. $Mg^{2\oplus}$ ions are important part of chlorophyll in green plants.

ii. $Mg^{2\oplus}$ ions play an important role in the breakge of glucose and fat molecules, in synthesis of proteins with enzymes, and in regulation of chlolesterol level.

iii. $Ca^{2\oplus}$ ions are important for bones and teeth in the form of apatite $[Ca_3(PO_4)_2]$

iv. $Ca^{2\oplus}$ ions play important role in blood clotting.

v. $Ca^{2\oplus}$ ions are required for contraction and stretching of muscles.

vi. $Ca^{2\oplus}$ ions are also required to maintain the regular beating of heart.

Problem 8.7 : Magnesium strip slowly tarnishes on keeping in air but metallic calcium is readily attacked by air. Explain.

Solution: Mg and Ca belong to group 2, but periods 2 and 3, respectively. During the reaction with air the metallic Mg and Ca lose their valence electrons. The tendency to lose valence electron is the metallic character, which increases down the group. Thus, calcium has higher metallic character, greater tendency to lose valence electron and lower ionization enthalpy than magnesium. Therefore Mg reacts slowly with air, forming a thin film of oxide, resulting into tarnishing, whereas Ca reacts readily at room temperature with oxygen and nitrogen in the air.

8.3 Some important compounds of elements of s-block

In this section we consider five important compounds of s-block elements with reference to their prepartion, properties and uses.

8.3.1. Sodium Carbonate (washing soda) $Na_2CO_3.\ 10H_2O$

Prepartion: Sodium Carbonate is commercially prepared by **Solvay process**. In the **first stage** of process CO_2 is passed into a concentrated solution of NaCl which is saturated with NH_3. Crystals of sodium bicarbonate separate as a result of the following reactions.

Reactions in the first stage:

$$2\,NH_3(aq) + H_2O + CO_2(g) \rightarrow (NH_4)_2\,CO_3(aq)$$

$$(NH_4)_2CO_3(aq) + H_2O + CO_2(aq) \rightarrow 2\,NH_4HCO_3(aq)$$

$$NH_4HCO_3(aq) + NaCl(aq) \rightarrow NH_4Cl(aq) + NaHCO_3(s)$$

Sodium bicarbonate has low solubility, and therefore its crystals precipitate out which are formed as a result of the double decomposition reaction between ammonium bicarbonate and sodium chloride.

In the **second stage** the separated crystals of sodium bicarbonate are heated to obtain sodium carbonate.

$$2\ NaHCO_3(s) \xrightarrow{\Delta} Na_2CO_3(s) + H_2O(g) + CO_2(g)$$

In this process the **recovery of ammonia** is done by treating the solution of NH_4Cl obtained with slaked lime, $Ca(OH)_2$. The byproduct of this reaction is calcium chloride.

$$2\ NH_4Cl(aq) + Ca(OH)_2(s) \longrightarrow 2\ NH_3(g) + CaCl_2(aq) + H_2O(l)$$

Do you know ?

Potassium carbonate can not be prepared by Solvay process because potassium hydrogen carbonate is highly water soluble and cannot be precipitated by reaction with potassium choride.

Properties : Sodium carborate (washing soda) is a white crystalline solid having the formula Na_2CO_3, $10H_2O$. It is highly soluble in water. On heating the decahyadrate loses water molecules to form monohydrate. On heating above 373 K temperature monohydrate further loses water and changes into white anhydrous powder called soda-ash.

$$Na_2CO_3.10H_2O(s) \xrightarrow{373\ K} Na_2CO_3.H_2O(s) + 9H_2O(g)$$

$$Na_2CO_3.H_2O(s) \xrightarrow{>373\ K} Na_2CO_3(s) + H_2O(g)$$

Aqueous solution of sodium carbonate is alkaline because of its hydrolysis by the following reaction:

$$Na_2CO_3 + H_2O \rightarrow NaHCO_3 + NaOH$$

Uses

i. The alkaline properties of sodium carbonate are responsible for emulsifying effect on grease and dirt. It is used as cleaning material.

ii. It is used to make hard water soft (as a water softener), as it precipitates out the soluble calcium and magnesium salts in hard water as carbonates. For example : $Ca(HCO_3)_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + 2NaHCO_3(aq)$

iii. It is used for commercial production of soap and caustic soda.

iv. It is an important laboratory reagent.

8.3.2 Sodium hydroxide (caustic soda) NaOH

Preparation : Sodium hydroxide is commercially obtained by the electrolysis of saturated aqueous solution of sodium chloride. Brine solution is subjected to electrolysis in Castner-Kellner cell. Mercury is used as cathode and carbon rod as anode. Metallic sodium liberated at the cathode forms sodium amalgam. Chlorine gas is evolved at the anode.

Cathode reaction: $Na^{\oplus} + e^{\ominus} \xrightarrow{Hg}$ Na-amalgam

Anode reaction: $Cl^{\ominus} \longrightarrow 1/2\ Cl_2 + e^{\ominus}$

Sodium hydroxide is obtained by treating sodium amalgam with water, when hydrogen gas is liberated.

$$2Na\text{-}Hg + 2H_2O \longrightarrow 2NaOH + 2Hg + H_2$$

Properties : Sodium hydroxide is a white deliquescent solid, having melting point 591 K. It is highly water soluble and gives a strongly alkaline solution. The surface of the solution absorbs atmospheric CO_2 to form Na_2CO_3.

Uses

i. Sodium hydroxide is used in purification of bauxite (the aluminium ore).

ii. It is used in commercial production of soap, paper, artificial silk and many chemicals.

iii. It is used for mercerising cotton fabrics.

iv. It is used in petroleum refining.

v. It is an important laboratory reagent.

8.3.3 Calcium Carbonate ($CaCO_3$)

Calcium carbonate is found in nature as chalk, lime stone, marble.

Preparation

i. When carbon dioxide is bubbled through solution of calcium hydroxide (slaked lime) water insoluble solid calcium carbonate is formed.

$$Ca(OH)_2(aq) + CO_2(g) \longrightarrow CaCO_3(s) + H_2O(l)$$

Excess carbon dioxide transforms the precipitate of $CaCO_3$ into water soluble calcium bicarbonate and therefore has to be avoided.

ii. When solution of calcium chloride is added to a solution of sodium carbonate, calcium carbonate is formed as precipitate.

$$CaCl_2(aq) + Na_2CO_3(aq) \longrightarrow CaCO_3(s) + 2NaCl(aq)$$

Properties

Calcium carbonate is soft, light, white powder. It is practically water insoluble. On heating to 1200 K calcium carbonate decomposes into calcium oxide and carbon dioxide.

$$CaCO_3(s) \xrightarrow{1200\ K} CaO(s) + CO_2\uparrow$$

It reacts with dilute acids to give the corresponding calcium salt and carbon dioxide.

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2\uparrow + H_2O$$

$$CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + CO_2\uparrow + H_2O$$

Uses:

i. Calcium carbonate in the form of marble is used as building material.

ii. Calcium carbonate is used in the manufacture of quicklime (CaO) which is the major ingredient of cement.

iii. A mixture of $CaCO_3$ and $MgCO_3$ is used as flux in the extraction of metals from ores.

iv. It is required for the manufacture of high quality paper.

v. It is an important ingredient in toothpaste, chewing gum, dietary suppliments of calcium and filler in cosmetics.

8.3.4 Hydrogen peroxide (H_2O_2) :

Hydrogen peroxide is a low cost, clean and mild oxidising agent. A 30% aqueous solution hydrogen peroxide is commercially available.

Preparation :

i. Hydrated barium peroxide is treated with ice cold dilute sulfuric acid. The precipitate of barium sulphate formed is filtered off to get hydrogen peroxide solution.

$$BaO_2.8H_2O(s) + H_2SO_4(aq) \xrightarrow[273\ K]{low\ temp} BaSO_4\downarrow + H_2O_2(aq) + 8H_2O(l)$$

ii. Small quantity of sodium peroxide is added to ice-cold solution of dilute sulfuric acid with stirring gives hydrogen peroxide (Merck process).

$$Na_2O_2\ (aq) + H_2SO_4\ (aq) \xrightarrow{273\ K} H_2O_2\ (aq) + Na_2SO_4\ (aq)$$

iii. A 50 % solution of sulfuric acid is subjected to an electrolytic oxidation to form peroxydisulfuric acid at anode.

$$2HSO_4^{\ominus} \xrightarrow{Electrolysis} H_2S_2O_8 + 2e^{\ominus}$$

Hydrolysis of the peroxydisulfuric acid yields hydrogen peroxide.

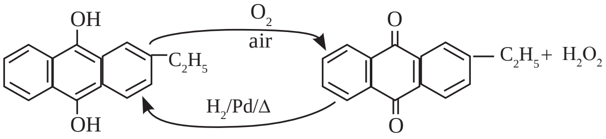
$$HO\text{-}SO_2\text{-}O\text{-}O\text{-}SO_2\text{-}OH + 2H_2O \rightarrow 2H_2SO_4 + H_2O_2$$

This method can be extended to laboratory preparation of D_2O_2

iv. Industrially hydrogen peroxide is prepared by air-oxidation of 2-ethylanthraquinol.

The 2-ethylanthraquinol is regenerated by catalytic hydrogenation of 2-ethylanthraquinone.

2-ethylanthraquinol $\xrightarrow[air]{O_2}$ H_2O_2 + 2 - ethylanthraquinone



Properties

i. Pure H_2O_2 is a very pale blue coloured liquid, having b.p. 272.4 K.

ii. H_2O_2 is miscible in water and forms a hydrate (H_2O_2. H_2O)

iii. Strength of aqueous solution of H_2O_2 is expressed in 'volume' units. The commercially marketed 30% (by mass) solution of H_2O_2 has volume strength of 100 volume. It means that 1 mL of 30% solution of H_2O_2 will give 100 mL oxygen at STP.

iv. H_2O_2 acts as a mild oxidising as well as reducing agent.

a. Oxidising action of H_2O_2 in acidic medium

$$2Fe^{2\oplus}(aq) + 2H^{\oplus}(aq) + H_2O_2(aq) \rightarrow 2Fe^{3\oplus}(aq) + 2H_2O(l)$$

b. Reducing action of H_2O_2 in acidic medium

$$2\,MnO_4^{\ominus} + 6H^{\oplus} + 5H_2O_2 \rightarrow 2Mn^{2\oplus} + 8H_2O + 5O_2$$

Uses

i. Hydrogen peroxide is used as mouthwash, germicide, mild antiseptic, preservative for milk and wine and bleaching agent for soft materials due to its mild oxidising property.

ii. Hydrogen peroxide, due to its reducing property, is used as an antichlor to remove excess chlorine from fabrics which have been bleached by chlorine.

iii. Now a days it is also used in environmental chemistry for pollution control, restoration of aerobic condition to sewage water.

Promblem 8.8 : Calculate % (by mass) of a H_2O_2 solution which is 45.4 volume.

Solution: 45.4 Volume H_2O_2 solution means 1L of this solution will give 45.4 L O_2 at STP

$2\,H_2O_2$	→	$2\,H_2O + O_2$
(2×34) g		22.7 L at STP

Thus, 22.7 L O_2 at STP is produced by 68 g H_2O_2

$\therefore$ 45.4 L O_2 at STP is produced by $\dfrac{68\times45.4}{22.7}$ = 136 g H_2O_2

$\therefore$ Strength this H_2O_2 solution = 136 g/L

$$= \frac{136\text{ g }H_2O_2}{1000\text{ g water}} = 13.6\%\text{ (by mass)}$$

8.3.5 Lithium aluminium hydride ($LiAlH_4$)

Lithium aluminium hydride is commonly abbreviated as LAH. It has chemical formula $LiAlH_4$.

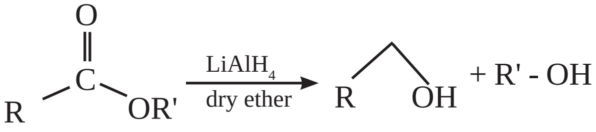
Prepartion : Lithium hydride is treated with aluminium chloride to give lithium aluminium hydride

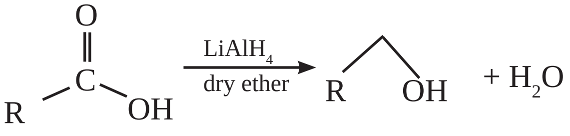
$$4LiH + AlCl_3 \rightarrow LiAlH_4 + 3LiCl$$

Properties: Lithium aluminium hydride is a colourless solid. It reacts violently with water and even atmospheric moisture.

Uses

(i) LAH is a source of hydride and therefore used as reducing agent in organic synthesis.





(ii) LAH is useful to prepare PH_3(phosphine)

$$4\,PCl_3 + 3LiAlH_4 \rightarrow 4\,PH_3 + AlCl_3 + Li\,Cl$$

Exercises

1. Explain the following

A. Hydrogen shows similarity with alkali metals as well as halogens.

B. Standard reduction potential of alkali metals have high negative values.

C. Alkaline earth metals have low values of electronegativity; which decrease down the group.

D. Sodium dissolves in liquid ammonia to form a solution which shows electrical conductivity.

E. $BeCl_2$ is covalent while $MgCl_2$ is ionic.

F. Lithium floats an water while sodium floats and catches fire when put in water.

2. Write balanced chemical equations for the following.

A. CO_2 is passed into concentrated solution of NaCl, which is saturated with NH_3.

B. A 50% solution of sulphuric acid is subjected to electrolyte oxidation and the product is hydrolysed.

C. Magnesium is heated in air.

D. Beryllium oxide is treated separately with aqueous HCl and aqueous NaOH solutions.

3. Answer the following questions

A. Describe the diagonal relationship between Li and Mg with the help of two illustrative properties.

B. Describe the industrial production of dihydrogen from steam. Also write the chemical reaction involved.

C. A water sample, which did not give lather with soap, was found to contain $Ca(HCO_3)_2$ and $Mg(HCO_3)_2$. Which chemical will make this water give lather with soap? Explain with the help of chemical reactions.

D. Name the isotopes of hydrogen. Write their atomic composition schematically and explain which of these is radioactive ?

4. Name the following

A. Alkali metal with smallest atom.

B. The most abundant element in the universe.

C. Radioactive alkali metal.

D. Ions having high concentration in cell sap.

E. A compound having hydrogen, aluminium and lithium as its constituent elements.

5. Choose the correct option.

A. The unstable isotope of hydrogen is .....

a. H-1	b. H-2
c. H-3	d. H-4

B. Identify the odd one.

a. Rb	b. Ra
c. Sr	d. Be

C. Which of the following is Lewis acid ?

a. $BaCl_2$	b. KCl
c. $BeCl_2$	d. LiCl

D. What happens when crystalline Na_2CO_3 is heated ?

a. releases CO_2

b. loses H_2O

c. decomposes into $NaHCO_3$

d. colour changes.

Activity :

1. **Collect the information of preparation of dihydrogen and make a chart.**
2. **Find out the s block elements compounds importance/uses.**