19

HYDROGEN AND s-BLOCK ELEMENTS

Hydrogen, alkali metals (like sodium and potassium) and alkaline earth metals (like magnesium and calcium) are the essential parts of the world we live in. For example, hydrogen is used in making vanaspati. Yellow glow of street light is due to sodium. Sodium choloride, potassium chloride and compounds of alkali metals are essential for life. Sodium hydroxide sold under the name of caustic soda is used in the manufacture of soap. Plaster of paris, a compound of calcium is used as a building material as well as by doctors in setting of bone fracture.

In this lesson we shall study occurrence, physical and chemical properties and uses of hydrogen and s-block elements (alkali metals and alkaline earth metals).

Objectives

After reading this lesson, you will be able to:

- explain the unique position of hydrogen in the periodic table;
- compare and contrast the properties of different isotopes of hydrogen;
- recall the various physical and chemical properties and uses of hydrogen with chemical reactions;
- explain the structure of water molecule and ice;
- list the uses of heavy water;
- list the different methods of preparation of hydrogen peroxide;
- list oxidizing and reducing properties of hydrogen peroxide with at least two examples of each;
- list the uses of hydrogen peroxide;
- recall the names and formulae of some common ores of alkali and alkaline earth metals;
- recall the electronic configuration of alkali and alkaline earth metals;

- write reactions of alkali and alkaline earth metals with oxygen, hydrogen, halogens and water;
- explain the trend of basic nature of oxides and hydroxides and
- explain the solubility and thermal stability of their carbonates and sulphates.

19.1 Hydrogen

Hydrogen is the first element of the periodic table. Hydrogen has the simplest atomic structure and consists of a nucleus containing one proton with a charge +1 and one orbital electron. The electronic structure may be written as $1s^1$.

19.1.1 Position in the Periodic Table

Where is hydrogen placed in periodic table?

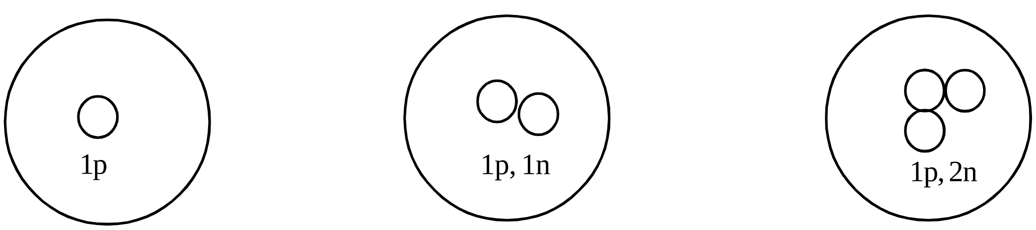
Elements are place in the periodic table according to their outermost electronic configuration. So hydrogen ($1s^1$) may be placed with alkali metals (ns^1). But hydrogen attains noble gas configuration of helium atom ($1s^2$) by gaining one electron. It forms the hydride ion H^- ($1s^2$) like halogens (ns^2np^5) by gaining one electron. On electrolysis of used alkali hydride, hydrogen is liberated at anode just as chlorine is liberated at anode during electrolysis of sodium chloride. Thus hydrogen ought to be placed in group 17 along with halogens. Hydrogen also resembles group 14 elements, since both have a half filled shell of electrons. So where should hydrogen be placed? This problem is solved by placing hydrogen neither with alkali metals nor with halogens. It has been given a unique position in the periodic table (see Periodic Table in lesson 4)..

19.1.2 Isotopes of hydrogen

If atoms of the same element have different mass numbers they are called isotopes. This difference in mass number arises because the nucleus contains a different number of neutrons.

Naturally occurring hydrogen contains three isotopes: protium ${}^1_1H$ or H, deuterium ${}^2_1H$ or D and tritium ${}^3_1H$ or T. These three isotopes contain one proton and 0, 1 and 2 neutrons, respectively in the nucleus (Fig 19.1). Protium is by far the most abundant.

Naturally occurring hydrogen contains 99.986% of the ${}^1_1H$ isotope, 0.014% of D and 7×10^{-16}% of T, therefore the properties of hydrogen are essentially those of the lightest isotope. Tritium is radioactive and emits low energy β particles ($t_{1/2}$ = 12.33yrs).



Protium	Deutirium	Tritium
Atomic number = 1	Atomic number = 1	Atomic number = 1
Mass number = 1	Mass number = 2	Mass number = 3
${}^1_1H$	${}^2_1H$ or D	${}^3_1H$ or T

Fig 19.1: *Isotopes to hydrogen*

Due to difference in mass of different isotopes, there arise a few differences in their properties. For example:

1. H_2 is more rapidly adsorbed on the metal surface than D_2.
2. H_2 reacts over 13-times faster with Cl_2 than does D_2.

Difference in properties that arises from the difference in mass is called isotope *effect*. Since the percentage difference in the mass of isotopes of hydrogen is very large. The difference in properties of isotopes of hydrogen is very large. The difference in properties of compounds containing these isotopes is also large.

19.1.3 Physical properties

Hydrogen is a diatomic gas, H_2. It is colourless and has no smell. It is lightest of all the gases known. It is insoluble in water, acids and most of the organic solvents. It is adsorbed when passed over platinum and palladium.

19.1.4 Chemical properties

1. Combustion: Hydrogen is combustible and burns in air with pale blue flame.

$$2H_2 + O_2 \rightarrow 2H_2O$$

2. Reducing property: Hydrogen reduces heated metal oxides to metals.

$$ZnO + H_2 \rightarrow Zn + H_2O$$

$$CuO + H_2 \rightarrow Cu + H_2O$$

3. Reaction with non-metals: Hydrogen combines with nitrogen, carbon, oxygen and chlorine under appropriate conditions to form ammonia, methane, water and hydrogen chloride, respectively.

$$3H_2 + N_2 \rightarrow 2NH_3$$

$$2H_2 + C \rightarrow CH_4$$

$$2H_2 + O_2 \rightarrow 2H_2O$$

$$H_2 + Cl_2 \rightarrow 2HCl$$

Reaction with metals: Hydrogen reacts with highly electropositive metals to from the corresponding hydrides.

$$2Na + H_2 \rightarrow 2NaH$$

$$2Li + H_2 \rightarrow 2LiH$$

19.1.5 Uses

Hydrogen is used:

1. for conversion of coal into synthetic petroleum.
2. in the manufacture of bulk organic chemicals, particularly methanol.
3. in the hydrogenation of oils. Vegetable oils change in to vegetable ghee when hydrogen is passed through the oils at 443K in presence of nickel as catalyst.

4. in the manufacture of ammonia, which is used in the production of fertilizers.
5. as primary fuel for heavy rockets.
6. for filling balloons.

Intext Questions 19.1

1. Name the isotopes of hydrogen?

   ....

2. Name the isotope of hydrogen which is radioactive?

   ....

3. Why is hydrogen used for filling balloons?

   ....

4. Which gas is produced, when hydrogen combines with carbon?

   ....

5. Name the gas, which is used for the production of fertilizers.

   ....

6. How are vegetable oils changed into vegetable ghee?

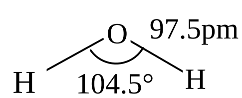
   ....

19.2 Compounds of Hydrogen

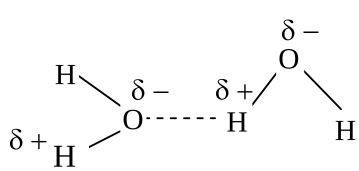
Hydrogen forms a large number of compounds: here we will consider only two of them, [illegible]$_2$O) and hydrogen peroxide (H_2O_2).

19.2.1 Water (H_2O)

This oxide of hydrogen is essential to all life. It occurs in the form of snow, as water in rivers, lakes, sea etc. and as vapour in the atmosphere. Water is a covalent compound made up of two hydrogen atoms linked with one oxygen atom through covalent bonds. Its Lewis structure and molecular structure are shown below.



Because of the large electronegativity of oxygen, water molecule is highly polar. It has partial negative charge (δ^-) on the oxygen atom and partial positive charge (δ^+) on the hydrogen atom. An electrostatic attraction between H of one molecule with O of other molecule results in the formation of intermolecular hydrogen bonds.



Structure of ice:

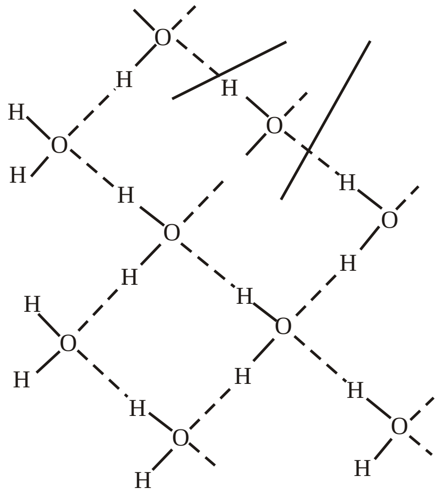


Fig. 19.2: *Tetrahedral arrangement of oxygen atoms in ice.*

The remarkable characteristic of water is that in solid form, it is less dense than liquid form. Consequently an ice cube floats on water. Water molecules are joined together in an extensive three dimensional network in which oxygen atom is bonded to four hydrogen atoms, two by hydrogen bonds and two by normal covalent bonds, in a near tetrahedral hydrogen bonded structure (Fig. 19.2), which has got open spaces. This is responsible for low density.

19.2.2 Heavy water and its applications

Water containing deuterium in place of ordinary hydrogen (protium) is termed as heavy water (D_2O). Heavy water is separated from water by electrolysis. The equilibrium constant for the dissociation of water containing protium is very high (1.0×10^{-14}) as compared to water containing deuterium (3.0×10^{-15}) e.g.

$$H_2O \rightarrow H^+ + OH^-$$

$$D_2O \rightarrow D^+ + OD^-$$

O–H bonds are broken more rapidly than O–D bonds. Thus when water is electrolyzed, H_2 is liberated much faster than D_2, and the remaining water thus becomes enriched in heavy water D_2O. In order to obtain one litre of almost pure D_2O, we have to electrolyze about 30000 litres of ordinary water.

Uses:

1. Heavy water is used as a moderator in nuclear reactors. In this process the high speed neutrons are passed through heavy water in order to slow down their speed.
2. It is used in the study of mechanism of chemical reactions involving hydrogen.
3. It is used as the starting material for the preparation of a number of deuterium compounds, for example:

$$CaC_2 + 2D_2O \longrightarrow C_2D_2 + Ca(OD)_2$$

$$SO_3 + D_2O \longrightarrow D_2SO_4$$

19.2.3 Hydrogen peroxide (H_2O_2)

Hydrogen peroxide is an important compound of hydrogen. Its chemical formula is H_2O_2.

Methods of preparation:

Two methods of preparation of hydrogen peroxide are given below:

1. By the action of dilute mineral acids (H_2SO_4) on metallic peroxides (barium peroxide, sodium peroxide)

$$BaO_2 . 8H_2O + H_2SO_4 \longrightarrow BaSO_4 + H_2O_2 + 8H_2O$$

$$Na_2O_2 + H_2SO_4 \rightarrow Na_2SO_4 + H_2O_2$$

2. By the electrolysis of H_2SO_4 (50% W/W) followed by distillation

At cathode: $2H^+ + 2e^- \rightarrow H_2$

At anode: $2SO_4^{2-} \rightarrow S_2O_8^{2-} + 2e^{-1}$

The anodic solution which contains persulphate ions $(S_2O_8)^{2-}$ is distilled with sulphuric acid at reduced pressure yielding H_2O_2:

$$S_2O_8^{2-} + 2H^+ + 2H_2O \longrightarrow 2H_2SO_4 + H_2O_2$$

Properties:

Hydrogen peroxide is a clolourless syrupy liquid and has sharp odour. It has a boiling point of 423K. It is miscible in all proportions with water, alcohol and ether. The oxidation state of oxygen in hydrogen peroxide is –1, a value, which lies between the oxidation state of oxygen in O_2 (zero) and water (–2). Therefore, hydrogen peroxide acts as an oxidizing agent as well as a reducing agent in acidic and alkaline media.

Oxiding Properties:

(a) Oxidizing action in acid solution:

(i) $PbS + 4H_2O_2 \rightarrow PbSO_4 + 4H_2O$

(ii) $2FeSO_4 + H_2SO_4 + H_2O_2 \rightarrow Fe_2(SO_4)_3 + 2H_2O$

(b) Oxidizing action in alkaline solution:

$MnCl_2 + H_2O_2 + 2KOH \rightarrow 2KCl + 2H_2O + MnO_2$

(c) Reducing action in acid solution:

$2KMnO_4 + 3H_2SO_4 + 5H_2O_2 \rightarrow 2MnSO_4 + K_2SO_4 + 8H_2O + 5O_2$

(d) Reducing action in alkaline solution:

(i) $2KMnO_4 + 3H_2O_2 \rightarrow 2MnO_2 + 3O_2 + 2H_2O + 2KOH$

(ii) $Cl_2 + H_2O_2 + 2KOH \rightarrow 2KCl + 2H_2O + O_2$

Uses:

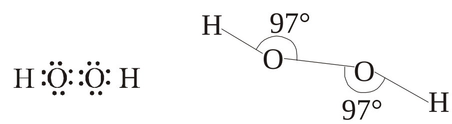
Hydrogen peroxide is used:

1. for bleaching hair, leather and wool etc.

2. as a germicide and disinfectant.
3. as an explosive when mixed with alcohol.
4. in the preparation of foam rubber.
5. in pollution control e.g. treatment of drainage and sewage water for dechlorination.

Structure:

The Lewis structure and molecular structure of hydrogen peroxide are shown below:



Intext Questions 19.2

1. Why does ice float on water?

   ....

2. What is heavy water? Write its important uses.

   ....

3. Give one method of preparation of hydrogen peroxide.

   ....

4. Give two uses of hydrogen peroxide.

   ....

5. How does hydrogen peroxide decolorize potassium permaganate?

   ....

19.3. *s*-Block Elements

The *s*-block elements have an outer electronic configuration ns^1 or ns^2 and are placed in the group 1 and 2 of the periodic table. Group 1 consists of the elements: lithium, sodium, potassium, rubidium, caesium and francium. They are collectively known as the alkali metals after the Arabic word *al-qis* meaning plant ashes. These ashes are particularly rich in carbonates of calcium, strontium, barium and radium. They are collectively known as alkaline earth metals.

19.3.1 The alkali metals

In this group all the elements are electropositive metals and there exists resemblance between the elements owing to their similar outer electron configuration. The occurrence and properties of alkali metals are discussed below:

Occurrence:

Sodium and potassium are abundant. Sodium is found as sodium chloride in the sea water and as sodium nitrate (Chile saltpeter) in the deserts of Chile. Potassium too, is found in sea water, and also as carnallite ($KCl.MgCl_2.6H_2O$). Lithium, rubidium and caesium occur in a few rare aluminosilicates. Francium is radioactive; its longest-lived isotope ^{223}Fr has a half life of only 21 minutes.

19.3.1.1 Electronic configuration

The alkali metals with their symbols, atomic numbers and electronic configurations are listed below in Table 19.1:

Table 19.1: Electronic configuration of alkali metals

Element	*Symbol*	*Atomic number*	*Electronic configuration*
Lithium	Li	3	$1s^2, 2s^1$
Sodium	Na	11	$1s^2, 2s^2p^6, 3s^1$
Potassium	K	19	$1s^2, 2s^2p^6, 3s^2p^6, 4s^1$
Rubidium	Rb	37	$1s^2, 2s^2p^6, 3s^2p^6d^{10}, 4s^2p^6, 5s^1$
Caesium	Cs	55	$1s^2, 2s^2p^6, 3s^2p^6d^{10}, 4s^2p^6d^{10}, 5s^25p^6, 6s^1$

19.3.1.2 Physical properties of Alkali Metals

Alkali metals are placed in group 1 of periodic table. They readily form unipositive ions. As we go down the group the alkali metals show steady increase in size due to the addition of a new shell at each step. The increase in size of the atoms or ions, directly influences the physical and chemical properties of the alkali metals. Some physical properties are given in Table 19.2.

Table 19.2: Physical properties of alkali metals

Symbol	Ionic Radius (pm)	First Ionization enthalpy (kJ mol^{-1})	Electro negativity	Density (g cm^{-3})	M.P. K	Electrode Potential (E° volts)
Li	76	520.1	1.0	0.54	454	–3.05
Na	102	495.7	0.9	0.97	371	–2.71
K	138	418.6	0.8	0.86	336	–2.83
Rb	152	402.9	0.8	1.53	312	–2.89
Cs	167	375.6	0.7	1.90	302	–2.93

The trends in physical properties are listed in Table 19.3.

Table 19.3: Trends in physical properties

No.	Characteristic	Trend
1.	Oxidation state	All elements show +1 oxidation state
2.	Atomic/ionic radii	Li < Na < K < Rb < Cs Atomic and ionic radii increases since number of shells increase as we go down the group.
3.	Ionization energy	Li > Na > K > Rb > Cs As the size increases it becomes easier to remove an electron from the outermost shell.
4.	Electronegativity	Li > Na > K > Rb > Cs The electropositive character increases due to decrease in ionization enthalpy therefore electronegativity decreases.
5.	Metallic character	Li < Na < K < Rb < Cs Metallic character increases as we go down the group due to increase in electropositive character.
6.	Density	Li < Na > K < Rb < Cs Generally density increases from Li to Cs as the atomic mass increases (exception K).
7.	Melting point & boiling points	Li > Na > K > Rb > Cs Decreases down the group because of increasing size and weak intermetallic bond.
8.	Flame coloration	They show characteristic colors in the flame. The outermost electron absorbs energy and is excited to a higher energy level. This absorbed energy is remitted when the electron comes back to ground state. The difference in energy falls in the visible range of radiation hence the colors are seen. Li – Crimson red; Na – Yellow; K – Pale violet; Rb – Violet; Cs – Violet

19.3.1.3 Chemical Properties

Alkali metals are the most reactive metals in the whole periodic table due to their ease in losing outermost electron hence getting oxidized easily. As the ease of losing electrons increases, the reactivity increases down the group.

(i) Oxides: All alkali metals form oxides, which are basic in nature. Lithium forms only one type of oxide, lithium monoxide Li_2O. Sodium peroxide Na_2O_2 is formed when sodium is heated with oxygen. Other metals of this group also form superoxides MO_2 on reaction with oxygen.

$$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$$

$$2Na(s) + O_2(g) \rightarrow Na_2O_2(s)$$

$$K(s) + O_2(g) \rightarrow KO_2(s)$$

The formation of a particular oxide is determined by the size of the metal ion. Tiny lithium ion is not able to come in contact with sufficient number of peroxo ions. However, the ions of potassium, rubidium, caesium are large enough to come in close contact with peroxo ions and form stable structures as superoxides.

(ii) Reactivity towards water: Although lithium has the most negative E°, its reaction with water is considerably less vigorous than that of sodium which has the least negative E° among the alkali metals (Table 19.2). The low reactivity of lithium is due to small size and high ionization enthalpy. All the metals of the group react with water explosively to form hydroxide and liberate hydrogen.

$$2M + 2H_2O \rightarrow 2M^+ + 2OH^- + H_2$$

Basic character of oxides and hydroxides: The basic character of oxides and hydroxides of alkali metals increases with the increase in size of metal ion. So, lithium oxide and hydroxide are least basic whereas, caesium oxide and hydroxide are most basic in nature.

(iii) Hydrides: The alkali metals react with hydrogen at about 637K to form hydrides (MH), where M stands for alkali metals.

$$2M + H_2 \rightarrow 2MH$$

(iv) Halides: Alkali metals react with halogens to form halides:

$$2M + X_2 \rightarrow 2MX (X = F, Cl, Br, I)$$

19.3.1.4 Stability and Solubility of Carbonates and Sulphates:

The carbonates and sulphates of alkali metals are generally soluble in water and thermally stable. The carbonates are highly stable to heat and melt without decomposing. As the electropositive character increases down the group, the stability of the carbonates increases. Carbonate of lithium is not so stable to heat due to the small size of lithium.

Intext Questions 19.3

1. Name the important ores of sodium.

   ....

2. Arrange the alkali metals in order of increasing ionization enthalpy.

   ....

3. Which of the alkali metals forms only monoxide?

   ....

4. Write down the chemical equation for the reaction of sodium with water.

   ....

5. What type of bond exists in the hydrides of alkali metals?

   ....

6. Name the element which forms (i) peroxide, (ii) superoxide.

....

19.4 The Alkaline Earth Metals

You have seen a gradual increase in the size of the alkali metals as we move down the group 1 of the periodic table. Identical observations may be made in the case of alkaline earth metals placed in group 2 of the periodic table. Some physical properties of the alkaline earth metals are given in Table 19.4.

Table 19.4: Physical properties of the alkaline earth metals

Symbol	Ionic Radius (pm)	First Ionization enthalpy (kJ mol^{-1})	Electro negativity	Density (g cm^{-3})	M.P. K	Electrode Potential (E°) volts
Be	89	899	1.5	1.85	1562	–1.70
Mg	136	737	1.2	1.74	924	–2.38
Ca	174	590	1.0	1.55	1124	–2.76
Sr	191	549	1.0	2.63	1062	–2.89
Ba	198	503	0.9	3.59	1002	–2.90

An alkaline earth metal atom is smaller in size compared to its adjacent alkali metal. This is due to the added proton in the nucleus, which exerts a pull on the electrons in an atom resulting in squeezing of the atom. This reduction in size shows higher control of the nucleus on the electrons in the shells.

The ease of losing electrons makes the alkaline earth metals good reducing agents. But this property is less prominent as compared to the corresponding alkali metals.

19.4.1 Occurrence

The alkaline earth metals are too reactive to occur native. Magnesium is the second most abundant metallic element in the sea, and it also occurs as carnallite ($KCl.MgCl_2.6H_2O$) in earth crust. Calcium occurs as calcium carbonate (marble, chalk etc) and with magnesium as dolomite ($CaCO_3.MgCO_3$). Other ores of calcium are anhydrite ($CaSO_4$) and gypsum ($CaSO_4.2H_2O$). Strontium and barium are rare and are found as carbonates and sulphates. Beryllium too is rare and is found as beryl ($Be_3Al_2(SiO_3)_6$).

19.4.2 Electronic Configuration

The electronic configurations of the alkaline earth metals are listed in Table 19.5.

Table 19.5: Electronic configuration

Element	*Symbol*	*Atomic number*	*Electronic configuration*
Beryllium	Be	4	$1s^2, 2s^2$
Magnesium	Mg	12	$1s^2, 2s^2p^6, 3s^2$
Calcium	Ca	20	$1s^2, 2s^2p^6, 3s^2p^6, 4s^2$
Strontium	Sr	38	$1s^2, 2s^2p^6, 3s^2p^6d^{10}, 4s^2p^6, 5s^2$
Barium	Ba	56	$1s^2, 2s^2p^6, 3s^2p^6d^{10}, 4s^2p^6d^{10}, 5s^25p^6, 6s^2$

19.4.3 Physical properties of alkaline earth metals

Alkaline earth metals are less electropositive than alkali metals. The electropositive character of alkaline earth metals increases down the group. They achieve an inert gas configuration by the loss of two electrons. Some physical properties and their trends are given in Table 19.6.

Table 19.6: Trends in physical properties

No.	Characteristic	Trend
1.	Oxidation state	All elements show +2 oxidation state
2.	Atomic/ionic	Be < Mg < Ca < Sr < Ba
	radii	Size of alkaline earth metals increases from top to bottom due to increase in the number of shells.
3.	Ionization	Be > Mg > Ca > Sr > Ba
	enthalpy	As the size increases it becomes easier to remove an electron from the outermost shell.
4.	Electronegativity	Be > Mg > Ca > Sr > Ba
		As the electropositive character increases from top to bottom due to decrease in ionization energy, electronegativity decreases from top to bottom.
5.	Metallic	Be < Mg < Ca < Sr < Ba
	character	Metallic character increases as we go down the group due to increase in electropositive character.
6.	Density	Generally density increases from top to bottom as the atomic mass increases.
7.	Melting point &	They show higher values of melting and boiling points as compared to
	boiling point	alkali metals because of the smaller size and stronger metallic bonds. There is no regular trend down the group. It depends upon packing.
8.	Flame coloration	Except Be and Mg (due to small size and higher ionization enthalpy) all other alkaline earth metals impart characteristic colours to the Bunsen flame.
		Ca — Brick red; Sr — Crimson red; Ba — Sea green

19.4.4 Chemical Properties of Alkaline Earth Metals

The alkaline earth metals are reactive metals, though less reactive than alkali metals. The reactivity increases from top to bottom in a group due to increase in electropositive character.

(i) Reactivity and E° values: The near constancy of the E° (M^{2+}/M) values for group 2 metals (Table 19.4) is somewhat similar to that for group 1 metals. Therefore, these metals are electropositive and are strong reducing agents. The less negative value for Be arises from, the large hydration energy associated with the small size of Be^{2+} being countered by relatively large value of the enthalpy of atomization of beryllium

(ii) Oxides: The alkaline earth metals burn in oxygen forming the ionic oxides of the type MO where M stands for alkaline earth metals except Sr, Ba, and Ra which form peroxides. Peroxides are formed with increasing ease and increasing stability as the metal ions become larger.

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

$$2Be + O_2 \rightarrow 2BeO$$

$$2Ca + O_2 \rightarrow 2CaO$$

$$Ba + O_2 \rightarrow 2BaO$$

Basic character of the oxides increases gradually from BeO to BaO. Beryllium oxide is amphoteric, MgO is weakly basic while CaO is more basic.

(iii) Hydrides: The alkaline earth metals combine with hydrogen to form hydrides of general formula MH_2

$$M + H_2 \rightarrow MH_2 (M = Mg, Ca, Sr, Ba)$$

(iv) Reaction with water: Usually the alkaline earth metals react with water to liberate hydrogen. Be does not react with water or steam even at red heat and does not get oxidized in air below 837K.

$$Mg + H_2O \rightarrow MgO + H_2$$

Ca, Sr, and Ba react with cold water with increasing vigour.

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

(v) Halides: All the alkaline earth metals combine directly with the halogens at appropriate temperature forming halides, MX_2 where M stands for alkaline earth metals.

$$M + X_2 \rightarrow MX_2$$

(vi) Solubility and stability of carbonates and sulphates:

Carbonates: The carbonates of alkaline earth metals are sparingly soluble in water. They decompose if heated strongly. Their thermal stability increases with increase in the size of the cation. Decomposition temperatures of carbonates are given below:

$BeCO_3$	$MgCO_3$	$CaCO_3$	$SrCO_3$	$BaCO_3$
<373K	813K	1173K	1563K	1633K

Sulphates: The sulphates of alkaline earth metals are white solids, stable to heat. The sulphates, $BeSO_4$ and $MgSO_4$ are readily soluble and the solubility decreases from $CaSO_4$ to $BaSO_4$. The greater hydration energies of Be^{2+} and Mg^{2+} ions overcome the lattice energy factor and therefore, their sulphates are soluble.

The sulphates decompose on heating, giving the oxides.

$MSO_4 \rightarrow MO + SO_3$

The thermal stability of sulphates increases with the increase in the size of cation.

This is shown by the temperature at which decomposition occurs:

$BeSO_4$	$MgSO_4$	$CaSO_4$	$SrSO_4$
773K	1168K	1422K	1647K

(vii) Complex compounds: Smaller ions of the group 2 elements form complexes. For example chlorophyll is a complex compound of magnesium. Beryllium forms complexes like $[BeF_4]^{2-}$.

Intext Questions 19.4

1. Name the common ore for K and Mg.

 ....

2. Arrange the alkaline earth metals in order of increasing reactivity.

 ....

3. Name an amphoteric oxide of alkaline earth metals.

 ....

4. Arrange the carbonates of alkaline earth metals in order of thermal stability.

 ....

What You Have Learnt

- Hydrogen can either be placed with alkali metals or with halogens.
- Hydrogen exists in three isotopic forms namely hydrogen, deuterium and tritium.
- Hydrogen is a combustible gas and has reducing property.
- There are two important oxides of hydrogen: water and hydrogen peroxide.
- Cage-like structure of ice makes it float on water.
- Water containing deuterium in place of ordinary hydrogen is known as heavy water.
- Heavy water can be separated from ordinary water by electrolysis or distillation.
- Heavy water is used as moderator in nuclear reactors.
- Hydrogen peroxide acts both as oxidizing and reducing agent.
- The alkali and alkaline earth metals show regular variation in various properties along a group and period.
- Alkali metals react with hydrogen, water and halogens to form hydrides, hydroxides and halides respectively.

- Basic nature of oxides and hydroxides of group 1 and group 2 elements.
- Thermal stability and solubility of carbonates and sulphates.

Terminal Exercise

1. Write three general characteristics of the s-block elements which distinguish them from the elements of other blocks.
2. The alkali metals follow the noble gases in their atomic structure. What properties of these metals can be predicted from this information?
3. What happens when?

   (a) sodium metal is dropped in water.

   (b) sodium metal is heated in free supply of air.

   (c) sodium peroxide dissolves in water.
4. Explain why hydrogen is best placed separately in the periodic table of elements.
5. Describe the industrial applications of hydrogen.
6. Discuss the importance of heavy water in nuclear reactor and how is it prepared from normal water?
7. Name the isotopes of hydrogen. What is the importance of heavier isotopes of hydrogen?
8. Why is ice less dense than water and what kind of attractive forces must be overcome to melt ice?
9. Show by proper chemical reactions how hydrogen peroxide can function both as an oxidizing and a reducing agent?
10. Compare the properties of alkali metals and alkaline earth metals with respect to:

    (a) atomic radii

    (b) ionization energy

    (c) melting points

    (d) reducing behavior
11. Explain the trends of solubility and stability of the carbonates and sulphates of alkaline earth metals.

Answers to Intext Questions

19.1

1. Three isotopes of hydrogen are (a) protium $^{1}_{1}H$, deuterium D or $^{2}_{1}H$ and (c) tritium T or $^{3}_{1}H$.

2. Tritium.
3. It is lightest of all the gases known.
4. Methane (CH_4).
5. Ammonia (NH_3).
6. Vegetable oils $+H_2 \xrightarrow[\text{Ni}]{443K}$ Vegetable ghee.

19.2

1. Ice is less dense as compared to water. It has open spaces in the hydrogen bonded structure.
2. D_2O; Moderator is nuclear reactors.
3. $BaO_2.8H_2O + H_2SO_4 \rightarrow BaSO_4 + H_2O_2 + 8H_2O$
4. (a) as a bleaching agent.

   (b) germicide and disinfectant.
5. H_2O_2 reduces $KMnO_4$

   $2KMnO_4 + 3H_2SO_4 + 5H_2O_2 \rightarrow 2MnSO_4 + K_2SO_4 + 8H_2O + 5O_2$

   Mn(+7) is reduced to Mn(+2)

19.3

1. NaCl and $NaNO_3$.
2. Cs < Rb < K < Na < Li
3. Lithium
4. $2Na + 2H_2O \rightarrow 2NaOH + H_2$
5. Ionic.
6. (i) Sodium (ii) potassium

19.4

1. Carnallite ($KCl.MgCl_2.6H_2O$).
2. Be < Mg < Ca < Sr < Ba
3. BeO
4. $BeCO_3 < MgCO_3 < CaCO_3 < SrCO_3 < BaCO_3$