

Chapter

1

PERIODIC CLASSIFICATION OF ELEMENTS AND PERIODICITY

INTRODUCTION:

“A table obtained by the arrangement of elements into periods and groups is called periodic table.”

The development of periodic table is one of the most significant achievement in the history of chemistry.

“The periodic table provides of basic framework to study the periodic behaviour of physical and chemical properties of elements as well as their compounds.”

HISTORICAL BACKGROUND

Al-Razi:

He organized the known elements and compounds. His classification was based upon the physical and chemical properties of elements and compounds.

Dobereiner (Law of Triads):

In 1829, Dobereiner discovered the families of elements which have same chemical properties. One family of elements consists of three elements, so he called it “triads”. In the given table each vertical column given is triad.

“The atomic mass of middle element is the average of atomic masses of other two elements of triads.”

Table – Some Doberiner's Triads

Element of triads with their atomic masses	Means of 1 st and third element
Li ₇ , Na ₂₃ , K ₃₉	$\frac{7 + 39}{2} = 23$
Cl _{35.5} , Br ₈₀ , I ₁₂₇	$\frac{35.5 + 127}{2} = 81$
Ca ₄₀ , Sr ₈₈ , Ba ₁₃₇	$\frac{40 + 137}{2} = 88.5$

NEWLAND'S LAW OF OCTAVE

In 1864, Newland, an English chemist classified 62 elements on the basis of their increasing atomic masses. He stated that:

“Every eight (8th) element had similar properties as first one, if they are arranged according to their atomic masses.”

Table – Newland's octaves

H	Li	Be	B	C	N	O
F	Na	Mg	Al	Si	P	S
Cl	K	Ca	Cr	Ti	Mn	Fe

DIMITRIE MENDELEEV'S PERIODIC TABLE

In 1871, a Russian chemist, Dimitrie Mendeleev gave a more useful and comprehensive scheme for the classification of elements. First of all, he arranged the elements into periods and groups. Vertical columns are called **groups** and horizontal rows are called **periods**. Mendeleev's periodic table consist of eight groups and twelve periods. He arranged the elements according to their ascending atomic masses and studied their behaviours.

“If elements are arranged according to their ascending atomic masses, he found that elements with similar chemical properties appeared at regular intervals.”

“The appearance of properties of elements after regular intervals is called periodic law or periodicity.”

Advantages of Mendeleev's Periodic Table:

Mendeleev left some gaps in his table for elements, which had not yet been discovered, and by considering their positions in the Periodic Table he predicted properties of these elements. For example, germanium was not known at that time but Mendeleev was confident that this element must exist so he predicted its properties. A few years later, germanium was discovered and its properties were same as predicted by Mendeleev.

Mendeleev corrected the atomic mass of Be. It was though to be 13.5 which was wrong. He proposed the atomic mass of Be to be 9.

Defects of Mendeleev's Periodic Table:

- (1) The basis of Mendeleev's periodic table was atomic mass. It did not give the idea of structure of atom.
- (2) Position of hydrogen could not be decided by Mendeleev.
- (3) Lanthanides (85 – 71) and Actinides (90 – 103) were not placed properly.
- (4) Position of isotopes was not indicated.
- (5) According to atomic masses, the position of the following pair of elements was not regular.

Ar	K	Co	Ni	Te	I
39.9	39	58.9	58.6	127.6	126.9

THE MODERN PERIODIC TABLE

"If elements are arranged according to their ascending order of atomic number, their chemical properties repeat in periodic manner."

In order to make the periodic table more useful and accurate, a few improvements were made by Mosley in the Mendeleev's periodic table.

Improvement in Mendeleev's Periodic Table:

- (1) After the discovery of x-rays, Mosley in 1913 discovered the atomic number of elements. *"Atomic number is the number of protons present in the nucleus of an atom."* According to Mosley, atomic number is the fundamental property of an element while atomic mass is not a fundamental property, as one element has different atomic masses for different isotopes. Hence periodic table was rearranged according to ascending order of **atomic numbers**. This improvement which was made by Mosley, was called modern periodic law. According to modern periodic law:

"If elements are arranged in the ascending order of their atomic numbers, their chemical properties repeat in a periodic manner."

- (2) Another improvement was the addition of an **extra group** (group VIII) at the extreme right of the Periodic Table. This group contains noble gases, which had not been discovered in Mendeleev's time.
- (3) Another confusion in Mendeleev's table was that elements like Be, Mg, Ca, Sr, Ba and Zn, Cd, Hg were placed in one vertical group, while according to their properties, they belonged to two different categories. The same was true for so many other elements placed in the same vertical group. In modern Periodic Table, the confusion was removed by introducing two types of vertical groups, A and B. In modern periodic table Be, Mg, Ca, Sr and Ba are placed in group IIA and Zn, Cd, Hg in group IIB.
- (4) Position of isotopes was adjusted because the periodic table was based on atomic number.
- (5) Position of Lanthanides and actinides was also adjusted.

[illegible]

**** Lanthanide series :**

+ Actinide series :

1. GROUPS AND PERIODS

(a) Groups:

"Elements with similar properties are placed in vertical columns called groups."

There are eight groups, which are usually numbered by **Roman numerals I to VIII**. Each group is divided into two sub-groups, designated as A and B sub-groups. The sub-groups, containing the typical or normal elements or representative elements, are labeled as A sub-groups whereas B sub-groups contain less typical elements, called transition elements and are arranged in the centre of the Periodic Table.

(b) Periods:

"The horizontal rows of the periodic table are called periods."

Period 1:

The period 1 contains only two elements, hydrogen and helium.

Period 2 and 3:

The period 2 and 3 contain eight elements each and are called **short periods**. All the elements in these periods are typical elements and belong to A sub-groups. In these periods, every eighth element resembles in properties with the first element. As lithium and beryllium in the 2nd period resemble in most of their properties with sodium and magnesium of the 3rd period respectively. Similarly boron and aluminium both show oxidation state, of +3, fluorine in 2nd period has close resemblances with chlorine of 3rd period.

Period 4 and 5:

The period 4 and 5 are called **long periods**. Each long period consists of eighteen elements. Out of these, eight are **typical elements** belonging to A sub-groups similar to second and third periods. Whereas the other ten elements, placed in the centre of the Table belong to B sub-groups and are known as **Transition Elements**. In these periods the repetition of properties among the elements occur after 18 elements. As after ${}_{19}\text{K}$ (having atomic number 19), the next element with similar properties is ${}_{37}\text{Rb}$.

Period 6:

The period 6 is also a **long period**, which contains thirty-two elements. In this periodic there are eight typical elements, ten transition elements and a new set of fourteen elements called **Lanthanides** as they start after ${}_{57}\text{La}$. Lanthanides have remarkably similar properties and are usually shown separately at the bottom of the Periodic Table.

Period 7:

The period 7 is incomplete so far. It contains only two normal elements ${}_{87}\text{Fr}$ and ${}_{88}\text{Ra}$, seven transition elements and fourteen inner transition elements. The inner transition elements of this period are called **Actinides** as they follow ${}_{89}\text{Ac}$. The actinides are also shown at the bottom of the Periodic Table under the Lanthanides. Due to their scarcity, the inner transition elements are also called **rare earth elements**.

Trans-uranic elements (elements after uranium) they are prepared in the laboratories.

2. SOME MORE FAMILIES

Elements of some groups have their collective names called family, e.g.,

Alkali Metals:

Elements of group IA are called alkali metals. Alkali metals are Li, Na, K, Rb, Cs and Fr.

Alkaline Earth Metals:

Elements of group IIA are called alkaline earth metals. They consist of Be, Mg, Ca, Sr, Ba and Ra.

Chalcogen Family:

Elements of group VIA are called chalcogen or ore forming group. Most of the ores contain oxygen or sulphur in them. They consist of O, S, Se, Te and Po.

Halogen Family:

The elements of group VIIA are called halogens or salt forming group. They consist of F, Cl, Br, I and At.

Noble Gases or Inert Gases:

Elements of group O or group VIIIA are called noble gases. Noble gases consist of He, Ne, Ar, Kr, Xe and Rn.

3. BLOCKS IN THE PERIOD TABLE

Elements in the Periodic Table can also be classified into **four blocks**. This classification is based upon the valence orbital of the element involved in chemical bonding. According to this classification:

- (1) **s-Block Elements:** The elements of IA and IIA sub-groups are called **s-block elements** because their valence electrons involve “s” orbital.
- (2) **p-Block Elements:** The elements of IIIA to VIIIA sub-groups (except He) are known as **p-block elements** as their valence electrons are present in “p” orbital.
- (3) **d-Block Elements:** As in transition elements, electrons in “d” orbital are responsible for their valence hence they are called **d-block elements**.
- (4) **f-Block Elements:** Lanthanides and Actinides, valence electrons are present in “f” orbital hence these elements are called **f-block elements**.

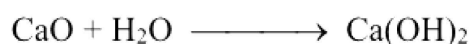
This classification is quite useful in understanding the chemistry of elements and predicting their properties, especially the concept of valency or oxidation state.

4. METALS, NON-METALS AND METALLOIDS

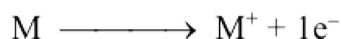
Elements of the periodic table can also be classified as metals, non-metals and metalloids.

Characters of Metals:

- (a) All metals have luster and good shine.
- (b) All metals are malleable (can be drawn into sheets) and ductile (can be drawn into wires).
- (c) All metals are good conductors of heat and electricity.
- (d) All metals form basic oxides, which give bases when dissolved in water e.g.,



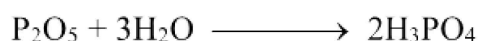
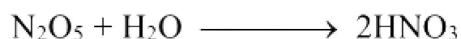
- (e) All metals have tendency to lose electrons easily and form positive ions.



Majority of the elements of periodic table are metals.

Characters of Non-metals:

- (a) Non-metals have no luster or shine.
- (b) Solid non-metals are brittle can be broken down into pieces by applying pressure in nature.
- (c) All non-metals are poor or non-conductors of heat and electricity.
- (d) All non-metals form acidic oxides, which give acids when dissolved in water and they react with bases to form salt and water.



- (e) Some non-metals gain electrons and form negative ions.

All the gases are non-metals. Element from group IVA to VIIIA top right of periodic table, above the stepped line are non-metals.

Characters of Metalloids or Semi-metals:

“The elements which have intermediate properties of metals and non-metals are called metalloids or semi-metals.”

Metalloids are located around the diagonal line running from boron to astatine except aluminium in the Periodic Table.

- (1) Oxides of the metalloids are amphoteric in nature.
- (2) They have less conductivity than metals and greater than non-metals.
- (3) Their conductivity increases by increasing temperature.

Metals				Metalloids	Non-metals				
Metals	Li	Be	d-block and f-block metals	B	C	N	O	F	Ne
	Na	Mg		Al	Si	P	S	Cl	Ar
	K	Ca		Ga	Ge	As	Se	Br	Kr
	Rb	Sr		In	Sn	Sb	Te	I	Xe
	Cs	Ba		Ti	Pb	Bi	Po	At	Rn
	Fr	Ra							

Classifying the element as metals metalloids and non-metals on the basis of their electrical conductance

Generally elements on the left hand side, in the center and at the bottom of periodic table are metals. Non-metal elements are present on the upper right of periodic table. Lower members of group VIIA, IVA and VA have the properties of both metals and non-metals. These elements are called semi-metals or metalloids. At the top right corner of group IVA to VIIIA, above the stepped line are non-metals. The elements just under the “steps” such as Si, As, and Te are metalloids. All the remaining elements except hydrogen are metals.

PERIODIC TREND IN PHYSICAL PROPERTIES

“The gradual change in the properties of elements when they are arranged according to their ascending order of atomic numbers is called periodic trend or periodicity.”

In periods, physical and chemical properties change gradually due to the change in atomic sizes and the electronic configuration of elements. In a group of the periodic table, chemical properties remain almost same but physical properties change due change in the atomic size.

1. ATOMIC SIZE

(a) Atomic Radius:

The average distance between the nucleus of the atom and its outer most shell provided the atom is assumed to be spherical.











































Size of a single atom cannot be measured directly because atoms have no sharp boundaries. Radius is defined as the distance from centre to boundary. We have studied that electrons have wave nature and concentration of cloud of electrons decreases rapidly

as the distance increases and it is not zero even at a large distance. It is impossible to know the boundary of an atom. Some techniques have been developed which can measure the distance between the nuclei of bonded atoms. Half of this distance is considered as atomic radius.

Trend of Atomic Radius in the Periodic Table:

In the periodic table, the atomic **radius increases** from top to bottom within a **group** due to increase in atomic number. This is because of the addition of an extra shell of electrons in each period. In a **period**, however, as the atomic number increases from left to right the atomic **radius decreases**. This gradual decrease in the radius is due to the increase in the positive charge in the nucleus. As the nucleus charge increases, electrons in shell are pulled closer to the nucleus. The overall size of the atom decreases. This effect is quite remarkable in the elements of longer periods in which 'd' and 'f' orbitals are involved. For example, gradual decrease in the size of lanthanides is called **Lanthanide contraction**.

Atomic radii of representative elements are given in the table.

I A	II A	III A	IV A	V A	VI A	VII A	VIII A
 H 37						 He 50	
 Li 152	 Be 111	 B 88	 C 77	 N 70	 O 66	 F 64	 Ne 70
 Na 186	 Mg 160	 Al 143	 Si 117	 P 110	 S 104	 Cl 99	 Ar 94
 K 231	 Ca 197	 Ga 122	 Ge 122	 As 121	 Se 117	 Br 114	 Kr 109
 Rb 244	 Sr 215	 In 162	 Sn 140	 Sb 141	 Te 137	 I 133	 Xe 130
 Cs 262	 Ba 217	 Tl 171	 Pb 175	 Bi 146	 Po 150	 At 140	 Rn 140

Atomic radii of the representative elements and the rare gases

(b) Ionic Radius:

“The radius of one of the ions in an ionic compound is called ionic radius. It is the measure of the spherical region around the nucleus within which electrons are most likely to be found.”

Size of Cation:

When a neutral atom loses one or more electrons, it becomes a positive ion or cation. The size of a **cation is smaller** as compared to its **neutral atom**. Decrease in size of cation is due to two reasons.

- (1) By removal of electrons usually number of shell decreases. e.g., Na has three shells while, Na^+ has only two shells in it.



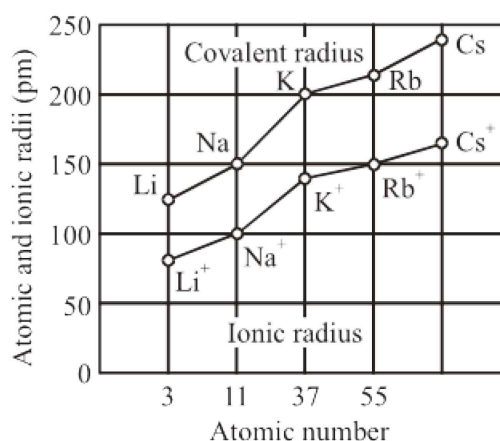
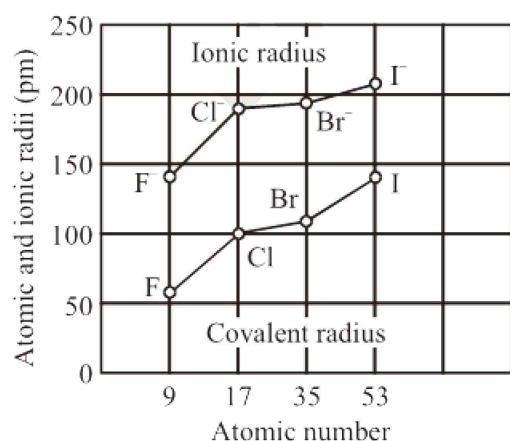
- (2) By removal of electrons, the ratio of electrons to protons, do not remain balance. By removal of electron, the nuclear charge increases and remaining electrons are drawn closer to the nucleus. e.g., The covalent radius of Na is 157 pm and ionic radius of Na^+ is 95 pm (Atomic radius = 186 pm).

Size of Anions:

When an atom gains electron, it changes to negative ion or anion. Size of anion is always larger than its neutral atom. By addition of more electrons, the nuclear charge decreases and shell expands. For example, the covalent radius of F is 72 pm while size of ionic F^- is 136 pm. The radius of Cl is 99 pm and that of Cl^- is 181 pm.

Trend of Ionic Radii in the Periodic Table:

In a group, the size of similar charged ions increases from top to bottom within a period. **Isoelectronic** (having same number of electrons) positive ions like Na^+ , Mg^{2+} , Al^{3+} show a decrease in the ionic radii from left to right due to increase in the nuclear charge. Same trend is observed in the negative ions. Size of anions with similar charge in a group increases from top to bottom. The size of isoelectronic negative ions (C^{4-} , N^{3-} , O^{2-} , F^-) decreases from left to right. The variation in atomic and ionic radii of alkali metals and halogens are shown below in the tables.



2. IONIZATION ENERGY

“The amount of energy which is required to remove an electron from the outermost shell of an isolated atom in gaseous state is called ionization energy.”

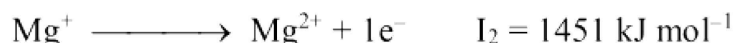
Ionization energy of sodium atom is 496 kJ mol^{-1} .



“The amount of energy which is required to remove first electron from atom in gaseous state is also called first ionization energy.” e.g.,



“The amount of energy required to remove second electron from monopositive ion is called second ionization energy.”



2nd I.E value is greater than 1st I.E value:

When an electron is removed from an atom, it changes to monopositive ion. The nuclear charge increases due to imbalance ratio of protons and electrons. Attraction of the nucleus increase on the remaining electrons of the cation. Thus for the removal of second electron, more amount of energy is needed than the first one.

TREND OF IONIZATION ENERGY VALUES IN PERIODIC TABLE

(a) Variation within the group:

The factor upon which the ionization energy of an atom mainly depends are:

- (1) Magnitude of nuclear charge
- (2) Size of the atom
- (3) Shielding effect.
- (4) Nature of the orbital.

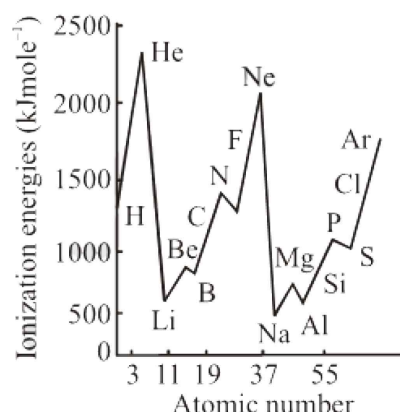
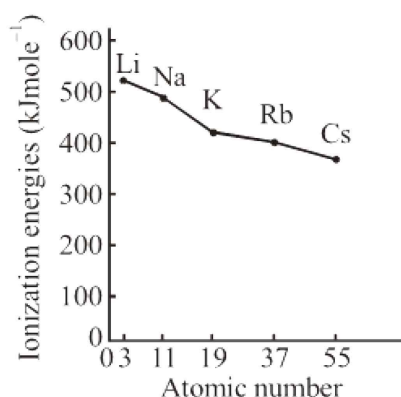
The shielding effect is actually the repulsion due to electrons in between the nucleus and the outermost shell. This effect increases, as the size of the atom increases due to addition of an extra shell successively in each period hence more number of electrons “shields” the nucleus. Going down in a group, the nuclear charge increases but as the size of the atom and the number of electrons causing the shielding effect also increases therefore ionization energy decreases from top to bottom. For example, it is easier to remove an electron from cesium atom than from lithium atom. The change in ionization energies of IA elements is shown below and also given in the form of table.

Table-I.E value in IA group in kJ per mol

Li	Na	K	Rb	Cs
520	496	419	403	376

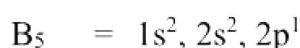
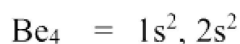
(b) Variation of I.E value across the period:

Generally, smaller the atom with greater nuclear charge, more strongly the electrons are bound to the nucleus and hence higher the ionization energy of the atom. By moving from **left to right in a period**, the outer shell remains the same while the nuclear charge increases effectively that makes the removal of an electron difficult and hence the value of ionization energy **increases**. Although the number of electrons also increases in this case but the shielding is not effective within the same shell. Therefore, due to complete outermost shell in them, the removal of electron is extremely difficult.

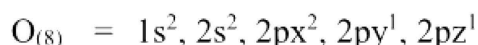
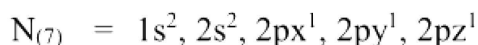
Table-I.E value of 2nd period (kJ per mole)

Li	Be	B	C	N	O	F	Ne
520	899	801	1086	1402	1314	1681	2080

In 2nd period I.E value generally increases from left to right along the period. Some abnormal values are also present between Be and B and N and O. Values of I.E for B and N decreases rather than increasing. This abnormal behaviour is due to stability of Be s-orbitals. Both s-orbitals are coupled.



Electronic configuration of nitrogen has three half filled p-orbitals which is more stable than the configuration of oxygen.



3. ELECTRON AFFINITY

“The amount of energy released or absorbed when an electron is added to an atom or gaseous ion to form gaseous negative ions.” e.g.,



Energy is usually released when electronegative elements absorb the first electron and E.A. in such cases is expressed in negative sign as in the case of halogens. When a **second electron** is added to a uninegative ion, the incoming electron is **repelled** by the already present negative charge and energy is absorbed in this process.



Unlike the ionization energy, electron affinities can be either positive or negative. Energy is required to add an electron to an already stable configuration. The electron affinity values for beryllium family elements, which have ns^2 configuration and for the noble gases, which have ns^2np^6 configuration are positive. All other E.A. values of representative elements are negative.

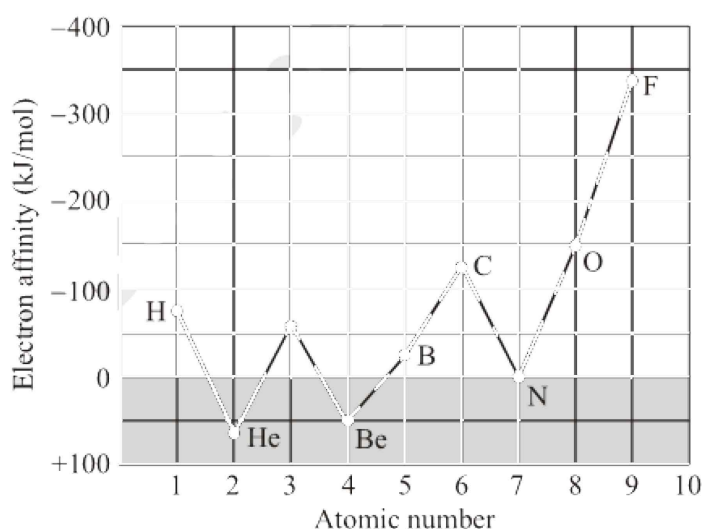
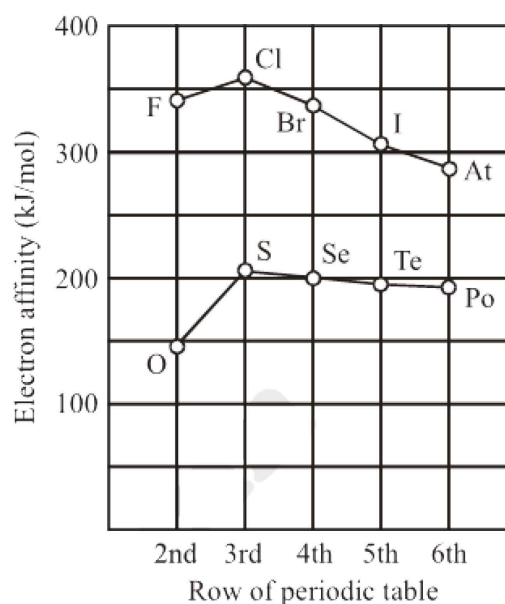
Group I							Noble gases
H -73	Group II	Group III	Group IV	Group V	Group VI	Group VII	He (21)
Li -60	Be (240)	B -83	C -123	N 0.0	O -141	F -337	Ne (29)
Na -53	Mg (230)	Al (-50)	Si -120	P -74	S -200	Cl -349	Ar (35)
K -48	Ca (156)	Ga (-36)	Ge -116	As -77	Se -195	Br -325	Kr (39)
Rb -47	Sr (168)	In -34	Sn -121	Sb -101	Te -183	I -295	Xe (41)
Cs -46	Ba (52)	Tl -50	Pb -101	Bi -101	Po (-170)	At (-270)	Rn (41)
Fr (-44)	Ra						

Electron affinities of the representative elements

TREND OF E.A IN PERIODIC TABLE

- (a) Electron affinity depends upon:
- (i) Size of atoms
 - (ii) Nuclear charge
 - (iii) Number of electrons in the valence shell.

The electron affinity generally **decreases** as we go down in a group due to two reasons. Firstly electrons are added to the larger orbital as compared to the first one. Secondly as we go down in a group force of attraction between the nucleus and electron decreases and they have less attraction. Fluorine and oxygen have less value of electron affinities due to smaller size and greater repulsion for incoming electron.



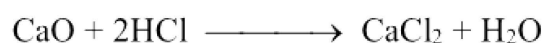
- (b) It appears that there is **no pattern** of electron affinity as shown in the graph. However they have relation with **electronic configuration**. Elements such as helium, beryllium, nitrogen and neon which have smaller affinity for extra electrons. Electronic configuration with either full filled or half filled sub-shells is stable.

4. METALLIC AND NON-METALLIC CHARACTERS

(a) Metallic Characters:

Elements of the periodic table can also be classified as metals, non-metals and metalloids. Metallic characters depend upon the size of atoms. Elements at the top right of periodic table are non-metals and the elements present at the left side are metals.

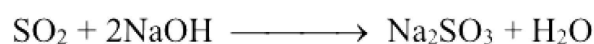
Chemically all the elements which have a tendency to form positive ions by **losing electrons** are considered **metals**. All metals are good **conductor** of heat and electricity, lustrous, malleable, ductile and forms **basic oxides**. Basic oxide when dissolved in water form hydroxides and they react with acids to form salt and water. e.g.,



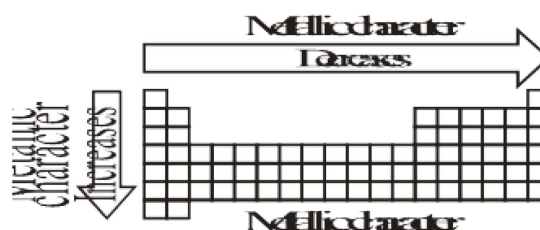
Metallic characters increases from top to bottom in groups due to increase in size of atom and ease to remove electrons. Metallic characters decreases from left to right in periods due to decrease in the atomic radii of elements. The elements of group VIIA (halogens) are least metallic in nature due to their smaller size.

(b) Non-metallic Characters:

The elements which gain electrons and form negative ions are called non-metals. All the gases are non-metals. Non-metals are poor conductor of heat and electricity, non-metals form acidic oxides. Acidic oxides when dissolved in water form acids. e.g., They and react with bases to form salt and water.



Non-metallic character of an element decreases as the atomic size increase. In a group of non-metals, like halogens, the non-metallic characters decreases from top to bottom.



Fluorine, which is present at the top is most non-metals. This trend can also be verified in the elements of group VA and VIA. Nitrogen and oxygen are pure non-metals and exist in gaseous state. While elements present at the bottom of these groups like, bismuth, polonium etc., are fairly metallic in nature.

5. MELTING AND BOILING POINTS

Melting and boiling points of elements explains that how strong the atoms or molecules are bonded together.

(a) Variation in Period:

Along the short periods, the melting and boiling points of elements **increase** with the number of valence electrons **upto group IVA** and then **decrease** upto the noble gases. The melting points of group IA elements are low because each atom in them provides only one electron to form a bond with other atom. Melting points of group IIA elements are considerably higher than those of group IA elements because each atom in them provides two binding electrons.

Since carbon has the maximum number of binding electrons, thus it has a very high melting point in diamond in which each carbon is bound to four other carbon atoms. In general, the elements which exist as **giant network covalent structures** have very high melting points.

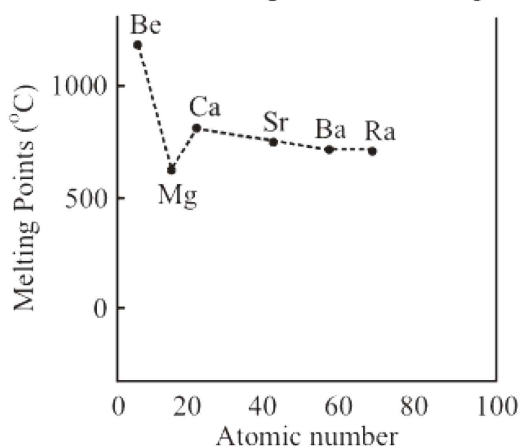
An important change occurs when we move from group IVA to VA, VIA, VIIA as the lighter elements of these groups exist as small, covalent molecules rather than as three dimensional lattices. For instance, nitrogen, oxygen and fluorine exist as individual diatomic molecules which have very weak intermolecular force between them. Consequently their melting and boiling points are extremely low.

(b) Variation in Groups:

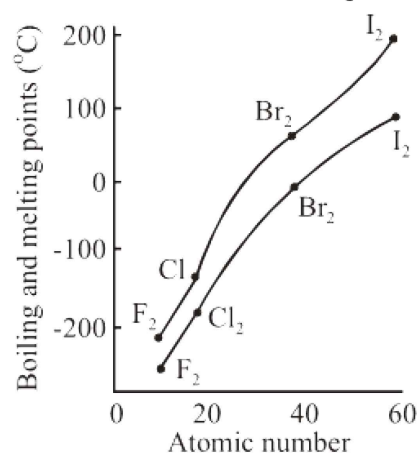
The melting and boiling points of **IA and IIA** group elements **decrease from top to bottom** due to the increase in their atomic sizes. The binding forces present between large sized atoms are relatively weaker as compared to those between smaller atoms.

For elements of **group VIIA** which exist in the form of molecules, the melting and boiling points **increase down the group**. This is because large molecules exert stronger force of attraction due to their higher polarizabilities and intermolecular forces.

F_2 and Cl_2 are gases, Br_2 is liquid and I_2 and At are solids at room temperature.



Melting points of Group IV elements



Melting and boiling point of halogens

6. OXIDATION STATE

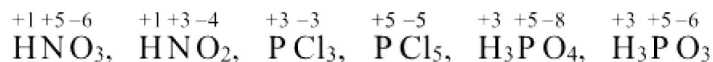
“The charge which is present on an element in a compound is called its oxidation state.”

- (1) In ionic compounds number of electrons which an atom loses or gains is its oxidation state. e.g., In NaCl, oxidation state of sodium and chlorine is +1 and -1 respectively.
- (2) In covalent compounds oxidation state of an element is decided by the difference of electronegativity between two atoms. Element with greater electronegativity has negative charge.

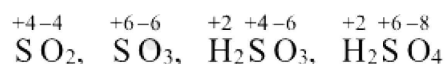
For example, SnCl_4 is a covalent compound. The oxidation state of Sn is +4 and chlorine is -1.

- (3) Oxidation state of an element is zero in its free state e.g., H_2^0 , Cl_2^0 , O_2^0 , N_2^0 , S^0 etc.
- (4) Element with maximum positive oxidation state of representative elements is equal to group number.

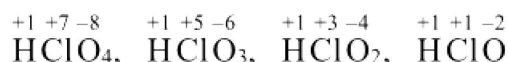
- (i) Elements of group IA and IIA have oxidation states +1 and +2 respectively.
- (ii) Elements of group IIIA B, Al, Ga etc. have oxidation state +3.
- (iii) Elements of group IVA C, Si, Ge, Sn, Pb have oxidation state +4.
- (iv) Elements of group VA N, P, As, Sb and Bi all shows oxidation state +3, and +5, +3 oxidation state is due to s-inert pair effect. e.g.,



- (v) Elements of group VIA show maximum oxidation +6 except oxygen. However, +4 oxidation state is also common. e.g.,



- (vi) Elements of group VIIA show maximum oxidation state, +7 except fluorine. +7, +5, +3 and +1. Oxidation states are more common is group VIIA. e.g.,



- (5) Metallic representative elements commonly show only positive oxidation state.
- (6) Most negative oxidation number found in representative elements is equal to the group number minus 8.

Negative oxidation states of group VA, VIA and group VIIA is -3, -2 and -1 respectively.

Group number $-8 =$ Negative oxidation state

Group $5 - 8 = -3$

Group $6 - 8 = -2$

Group $7 - 8 = -1$

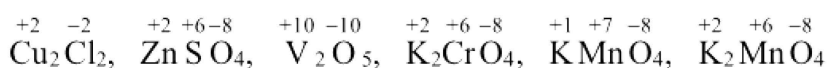
Negative oxidation state of an element is the number of electrons needed to complete the octet of valence shell.

- (7) Oxidation state of noble gases is zero because their outer most shell is complete.

**Table – Common oxidation number of representative elements
(only first three periods)**

Period	IA	IIA	IIIA	IVA	VA	VIA	VIIA	0
1	H +1 -1							He 0
2	Li +1	Be +2	B +3	C +4 to -4	N +5 to +1 -3	O -1 -2	F -1	Ne 0
3	Na +1	Mg +2	Al +3	Si +4 -4	P +5 +3 -3	S +6 +4 -2	Cl +7 +5 +3 +1 -1	Ar 0

Transition elements, which are shown in B sub-groups of the periodic table, also show the oxidation states equal to their group number as it can be seen for Cu(I), Zn(II), V(V) and Mn(VII). But due to greater number of valence electrons available in partially filled d-orbitals, these elements usually show more than one oxidation states in their compounds. For example:



7. ELECTRICAL CONDUCTANCE

“The capability of a substance to conduct electricity without change in composition is called conductivity.”

Electrical conductivity of metals depend upon the movement of free electrons in solid lattice. Following trend is seen in the periodic table about conductivity.

- (1) General trend in the period table is that, electrical conductivity is high for metals on left side of each period, lower for metalloids for the centre and negligible for the non-metals of each period.
- (2) The electrical conductance of metals in **group IA and IIA** generally **increases** from top to bottom. However the trend is not free from the individual variation in different atoms. Metals of group IB, which are known as coinage metals, have extraordinary high value of electrical conductance.
- (3) Non-metals, on the other hand, especially of groups VIA and VIIA, show such low electrical conductance that they can be considered as non-conductors.
- (4) In the series of transition metals the values of electrical conductance vary so abruptly that no general trend can be assigned to them.
- (5) Carbon, in the form of diamond is non-conductor because all of its valence electrons are tetrahedrally bound and unable to move freely, while in the form of graphite, carbon is fairly good conductor because one of its four valence electrons is relatively free to move.
- (6) The lower elements of group IVA, tin and lead, are fairly good conductors and their values of electrical conductivity are comparable with those of their counterparts in group IA.

8. HYDRATION ENERGY

“The amount of heat evolved or absorbed when one mole of gaseous ions or other solute particles dissolves in water to give an infinitely dilute solution.”

For example, when one mole of gaseous hydrogen ions are dissolved in water, resulting an infinitely dilute solution, a large amount of heat is liberated.



Hydration energies of a few negative and positive ions are shown in the table.

Table – Hydration energies of ions

Ion	$\Delta H_h \text{ kJ mol}^{-1}$	Ion	$\Delta H_h \text{ kJ mol}^{-1}$	Ion	$\Delta H_h \text{ kJ mol}^{-1}$	Ion	$\Delta H_h \text{ kJ mol}^{-1}$
Li^+	–499					F^-	–457
Na^+	–390	Mg^+	–1891	Al^{+++}	–4613	Cl^-	–384
K^+	–305	Ca^{++}	–1562			Br^-	–351
						I^-	–307

Variation in Group:

Hydration energy depends upon the **charge to size** or volume ratio of the ions. In **IA group**, the size of ions increases but charge is same, so the **hydration energy decreases**. Hydration energy of VIIA group also decreases downward due to decrease in the ratio of charge to size.

Variation in Period:

Hydration energy **in period increases** from left to right. For example, hydration energy of Mg^{2+} ion is greater than Na^+ due to greater ratio of charge to size.

Hydration energy of Al^{3+} is greater than Mg^{2+} also due to greater ratio of charge to size.

9. PERIODIC RELATIONSHIP IN COMPOUNDS

(a) Halides:

“Binary compounds of halogens with other elements are called halides.”

There are three types of halides:

- (i) Ionic halides (ii) Polymeric halides (iii) Covalent halides

(i) Ionic Halides:

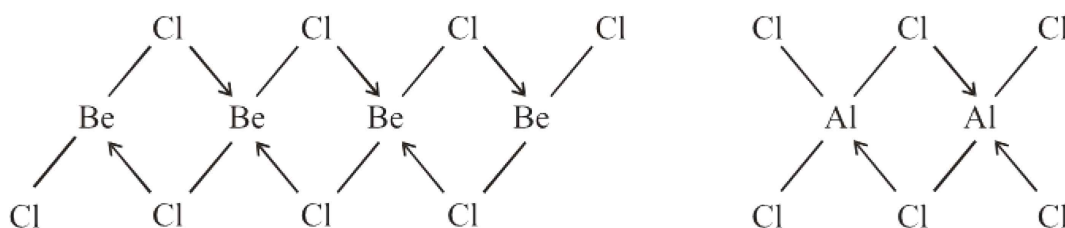
“If ionic bond is present between the metal and halogen, the halides are called ionic halides.”

All elements of group **IA form ionic halides** with halogens. Element of **group IIA** also form **ionic halides** with halogen except beryllium. Strongly electropositive elements, having greater electronegativity difference with halogen atom, form ionic halides. The halides of group IA are considered purely ionic compounds which have high melting points and solid in nature. Such halides have three dimensional lattices consist of discrete ions. Among the pure ionic compounds, the fluorides have the highest lattices energies due to the small size of fluoride ion. Thus for ionic halides the fluorides have the highest melting and boiling points which decrease in the order:

Fluoride > Chloride > Bromide > Iodide

(ii) Polymeric Halides:

“The halides which have polymeric or bridge structure are called polymeric halides.”



Usually electron deficient halides form such polymeric halides to complete their octet. Less electropositive elements such as Be, Ga and Al form polymeric halides having partly ionic bonding with layer or chain lattices.

(iii) Covalent Halides:

“The halides in which halogens form covalent bond with other elements are called covalent halides.”

Elements with high electronegativity form covalent halides. Elements of group IVA, VA and VIA form covalent halides. Some metalloids also form covalent halides. Some characters of covalent halides are given below.

(a) Bond Polarity and Melting Points of Halides:

The lattice of SiCl_4 consists of discrete molecules, which are highly polar. The bond in PCl_3 and S_2Cl_2 are less polar than those of SiCl_4 . On moving from left to right, the electronegativity difference reduces and the trend shifts towards covalent halides. The gradual change in bond type and melting points of the chlorides on moving across period 3 of the periodic table is shown in the following table.

Table – Melting points of chloride of period 3 elements and their bonding character

Property	NaCl	MgCl ₂	AlCl ₃	SiCl ₄	PCl ₃	S ₂ Cl ₂
M.Point (°C)	801	715	192	–68	–93	–80
Bonding	Ionic	Partly ionic	Partly ionic	Polar covalent	Polar covalent	Polar covalent

(b) Intermolecular Forces and Physical State of Halides:

As the intermolecular forces in covalent halide molecules are weak van der Waals' forces so they are often gases, liquids or low melting point solids. Physical properties of covalent halides are influenced by the size and polarizability of the halogen atom. **Iodides**, as being the largest and **more polarizable** ions possess the strongest van der Waals' forces and therefore have higher melting and boiling points than those of other covalent halides.

(c) *Nature of Halogens and Ionic Character of Halides:*

The variation in bonding character is also present in descending from top to bottom in the halogen group. In general, for a metal, the order of decreasing ionic character of the halides is: fluoride > chloride > bromide > iodide. For example AlF_3 is purely ionic compound having melting point 1290°C and fairly a good conductor whereas AlI_3 is predominantly covalent with melting point 198°C and electrically a non-conductor.

(d) *Oxidation State and Ionic Characters of Halides:*

In case of an element forming more than one halides, the metal halide in its **lower oxidation state tends to be ionic** while that in the higher oxidation state is covalent. For example, PbCl_2 is mainly ionic and PbCl_4 is fairly covalent. This can be explained by the high polarizing power of Pb^{+4} as compared to that of Pb^{+2} .

(b) **Hydrides:**

"The binary compounds of hydrogen with other elements are called hydrides."

According to the nature of bonding, hydrides may be broadly classified into three classes: ionic, covalent and intermediate.

Ionic Hydrides:

The elements of **group IA** and the **heavier members of group IIA** form ionic hydrides which contain H^- (hydride) ion. These hydrides are crystalline solid compounds, with high melting and boiling points and which conduct electricity in molten state. These hydrides are basic in nature. The tendency towards covalent character increases from left to right in the periodic table.

Intermediate Hydrides:

Hydrides of beryllium and magnesium represent the class of intermediate hydrides. Their properties are in between the ionic and covalent hydrides. They have polymeric structures and covalent nature.

Table – Hydrides of the elements of IA to VIIA and IIB sub-groups

IA	IIA	IIB	IIIA	IVA	VA	VIA	VIIA
LiH	BeH ₂		BH ₃	CH ₄	NH ₃	H ₂ O	HF
NaH	MgH ₂		AlH ₃	SiH ₄	PH ₃	H ₂ S	HCl
KH	CaH ₂	ZnH ₂	GaH ₃	GeH ₄	AsH ₃	H ₂ Se	HBr
RbH	SrH ₂	CdH ₂	InH ₃	SnH ₄	SbH ₃	H ₂ Te	HI
CsH	BaH ₂			PbH ₄	BiH ₃		
Ionic		Intermediate		Covalent			

Characters of Covalent Hydrides:

- (1) The covalent hydrides are usually gases or volatile liquids.
- (2) They are usually non-conductors and dissolve in organic solvents.
- (3) Their bond energies depend on the size and the electronegativity of the element.
- (4) Stability of covalent hydrides increases from left to right in a period and decreases from top to bottom in a group. Fluorine forms the most stable hydride and the least stable are those of thallium, lead and bismuth.
- (5) These hydrides are formed by elements with electronegativity value greater than 1.8 (Pauling scale). Since the electronegativity of hydrogen is 2.1, most of these hydrides have polar covalent bonds in which hydrogen is carrying a slight positive charge.
- (6) On moving from left to right along a period the electronegativity of the other element increases and the **hydrogen-element bond** becomes more polar. Due to high polarity the hydrides H_2O and HF are capable of forming hydrogen bonds between their molecules.
- (7) The **boiling points** of covalent hydrides generally **increase** on descending a group as shown in the table, except the hydrides like H_2O , HF and NH_3 which due to hydrogen bonding have higher boiling points than might be expected.
- (8) **Covalent hydrides are acidic** in nature except CH_4 which is neutral and NH_3 which is basic.

Table – Melting and boiling points of hydrides of groups IVA and VIA

Hydrides (Group IVA)	CH_4	SiH_4	GeH_4	SnH_4
M.P. ($^{\circ}\text{C}$)	-184	-185	-165	-150
B.P. ($^{\circ}\text{C}$)	-164	-112	-90	-52

Hydrides (Group VIA)	H_2O	H_2S	H_2Se	H_2Te
M.P. ($^{\circ}\text{C}$)	0.00	-82.9	-65.7	-48
B.P. ($^{\circ}\text{C}$)	100	-59.6	-41.3	-1.8

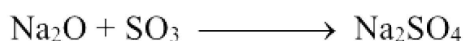
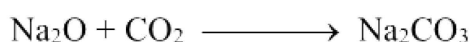
(c) Oxides:

"The compounds of oxygen with other elements are called oxides."

Oxides of representative elements are classified in many ways e.g.,

- (i) Types of bond present in oxides.
- (ii) Acidic or basic nature of oxides.
- (iii) Oxidation state of oxygen in oxides.

We know that metal forms basic oxides and non-metal forms acidic oxides. These acidic and basic oxides combine to form salts. For example:



"The oxides which show both acidic as well as basic properties are called amphoteric oxides."

Basic Oxides:

Oxides of **IA and IIA group** elements form **basic oxides** except the oxide of beryllium which forms intermediate or amphoteric oxide. Oxide of d-block transition elements are also basic but mostly insoluble in water. Oxide of IA and IIA have high affinity for proton and cannot exist alone in aqueous solution. Therefore, they accept proton and change to hydroxide ions OH^- .

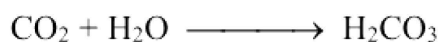
**Amphoteric Oxides:**

Oxides of relatively less electropositive elements are amphoteric in nature. For example, BeO , Al_2O_3 , ZnO , PbO etc., are amphoteric oxides. Amphoteric oxides behave as acid when react with strong base and behave as base when react with strong acids. e.g.,

- (i)
$$\text{ZnO} + \text{H}_2\text{SO}_4 \longrightarrow \text{ZnSO}_4 + \text{H}_2\text{O}$$
$$\text{ZnO} + 2\text{NaOH} + \text{H}_2\text{O} \longrightarrow \text{Na}_2[\text{Zn(OH)}_4]$$
- (ii)
$$\text{Al}_2\text{O}_3 + 6\text{HCl} \longrightarrow 2\text{AlCl}_3 + 3\text{H}_2\text{O}$$
$$\text{Al}_2\text{O}_3 + 2\text{NaOH} \longrightarrow 2\text{NaAlO}_2 + \text{H}_2\text{O}$$

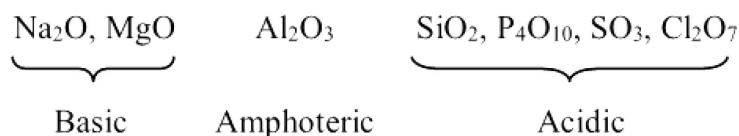
Acidic Oxides:

Oxides of non-metals such as C, N, P, S, Cl, F, Br etc. are acidic. When acidic oxides are dissolved in water, they form acids. e.g.,



Trend in Periodic Table:

In a period, acid-base characters change from basic (left hand side) to amphoteric and then acidic at right hand side. Example of oxides of third period is given below:



The basic characters of IA and IIA metal oxides increase downward in a group $\text{BeO} < \text{MgO} < \text{SrO} < \text{BaO}$.

In oxides of IIA group, BeO is less basic while BaO is more basic. The same trend is found in transition metal oxides. In IIB group ZnO is amphoteric while Cd and Hg oxides are basic.

Oxidation State of Metal and Nature of Oxides:

Acid-base characters of oxides also depend upon the oxidation state of metal. The basic characters decrease and **acidic characters increase by increasing the oxidation state** of metal ions. For example, in MnO , Mn_2O_3 , MnO_2 and Mn_2O_7 , the oxidation state of manganese is +2, +3, +4 and +7 respectively. The acidic character increases and basic character decreases as $\text{MnO} < \text{Mn}_2\text{O}_3 < \text{MnO}_2 < \text{Mn}_2\text{O}_7$. In this way Fe_2O_3 is less basic than FeO .

Table – Classification of oxide based on acid-base characters

Group	IA	IIA	IIB	IIIA	IVA	VA	VIA	VIIA
Valence shell configuration	s^1	s^2	s^2	s^2p^1	s^2p^2	s^2p^3	s^2p^4	s^2p^5
General formula	$(\text{M}^+)_2\text{O}^{2-}$	$\text{M}^{+2}\text{O}^{2-}$	$\text{M}^{+2}\text{O}^{2-}$	M_2O_3	MO MO_2	M_2O_3 M_2O_6	MO_2 MO_3	M_2O_5 M_2O_7

Li	Be	–	B	C	N	–	F
Na	Mg	–	Al	Si	P	S	Cl
K	Ca	Zn	Ga	Ge	As	Se	Br
Rb	Sr	Cd	In	Sn	Sb	Te	I
Cs	Ba	Hg	Tl	Pb	Bi	Po	At
Basic				Amphoteric		Acidic	

Nitrogen forms other oxides like N_2O , NO , NO_2 and N_2O_4 fluorine form OF_2 , O_2F_2 , but not F_2O_5 or F_2O_7 .

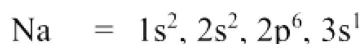
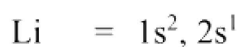
Bromine form various oxides as Br_2O and BrO_2 but Br_2O_3 and Br_2O_5 are not known.

POSITION OF HYDROGEN IN THE PERIODIC TABLE

Hydrogen is not a metal but in most of versions of the periodic tables, hydrogen is present at the top of group IA. Hydrogen has some similarities and dissimilarities with group IA, IVA and VIIA.

(a) Similarities of Hydrogen with Group IA:

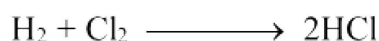
- (1) Hydrogen has one electron in its outermost shell and its electronic configuration is $1s^1$. Alkali metals also have electronic configuration ns^1 , e.g.,



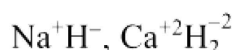
- (2) Hydrogen loses one electron like alkali metals and form H^+ .



- (3) Both hydrogen and alkali metals have strong tendency to combine with highly electronegative elements such as halogens.



- (4) Like alkali metals, hydrogen can also form ionic bond by transfer of electrons. These ionic compounds dissociate in water. e.g.,

**Dissimilarities of Hydrogen with Alkali Metals:**

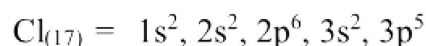
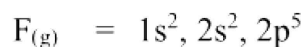
- (1) Hydrogen is a gas whereas alkali metals are solids (Cesium and Francium are liquids above 30°C).
- (2) Hydrogen needs only one electron to complete its valence shell whereas alkali metals need seven electrons to complete the outermost shell.
- (3) H^+ ion cannot exist independently in aqueous solution unlike alkali metals ions, e.g., Na^+ , K^+ etc., H^+ associates with H_2O to form H_3O^+ .



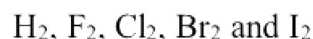
- (4) Alkali metals do not exist in diatomic state like H_2 .
- (5) Hydrogen is non-metal while the elements of IA group are metals.

(b) Similarities of Hydrogen with Halogens:

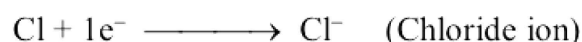
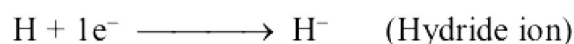
- (1) Hydrogen requires one electron to complete its outermost shell while elements of group VIIA also require one electron to attain the nearest noble gas configuration.



- (2) Hydrogen is a gas like most of the halogens.
 (3) Hydrogen and halogens exist in diatomic state. e.g.,



- (4) By accepting one electron hydrogen and halogens form negative ions.



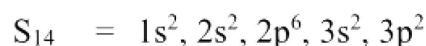
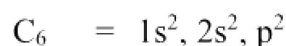
- (5) Both hydrogen and halogens form stable ionic compounds with alkali metals. e.g., NaH, NaCl.

Dissimilarities of Hydrogen with Halogens:

- (1) Hydrogen has one electron in the valence shell and halogens have seven electrons in the valence shell.
 (2) By losing electron, hydrogen form H^+ ions while halogens do not form positive ions.
 (3) Hydrogen combine with oxygen to form very stable oxide (H_2O) while halogen does not form such stable oxides.
 (4) Halide ions (\bar{F} , \bar{Cl} , \bar{Br} , \bar{I}) are stable in aqueous form while H^- is unstable in water.

**(c) Similarities of Hydrogen with Group IVA (Carbon Family):**

- (1) Valence shell of hydrogen is filled like those of group IVA, outer most shell of both hydrogen and carbon family is half filled.

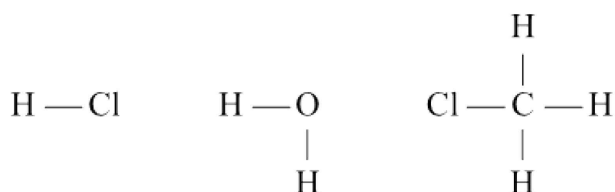


- (2) Hydrogen and carbon combine with other elements with covalent bond.
 (3) In majorities of reactions of carbon and hydrogen, they act as reducing agents.



Dissimilarities of Hydrogen with Group IVA:

- (1) Carbon and silicon combine with each other or with other elements to form chain type compounds while hydrogen does not show this property.
- (2) Hydrogen requires one electron to complete its outermost shell while elements of group IVA require four electrons.
- (3) Carbon can form bond with more than one elements at a time but hydrogen can only form one bond with an element.



Hydrogen exist in diatomic state (H_2) while elements of group IVA do not exist in diatomic state.

- (4) Hydrogen is gas while group IVA elements are all solids.
- (5) Hydrogen is s-block while carbon in p-block element.

Some of the properties of hydrogen are similar to those of the elements of certain groups, as discuss above, but this is a fact that hydrogen is a unique element whose properties do not match exactly with any of the groups in the periodic table. However due to partial resemblance in properties with alkali metals and monovalent nature, hydrogen is usually placed at the top of elements in group IA.