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 1st 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 (8^{th}) 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."

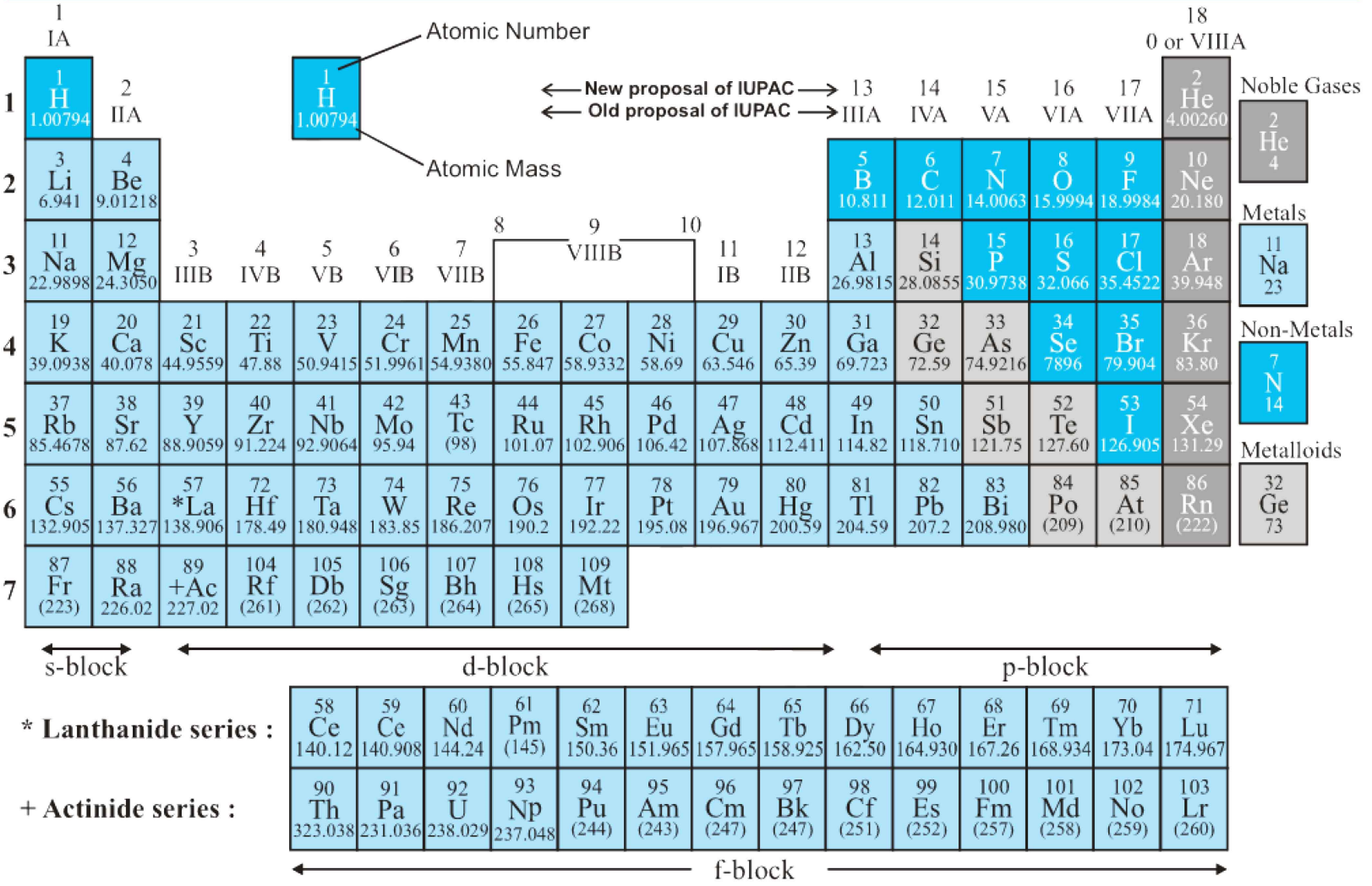
(2) Another improvement was the addition of an **extra group** (group VIII) at the extreme right of the Periodic Table. This group contains noble gases, which had not been discovered in Mendeleev's time.

(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.

MODERN PERIODIC TABLE OF THE ELEMENTS



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 2^{nd} period resemble in most of their properties with sodium and magnesium of the 3^{rd} period respectively. Similarly boron and aluminium both show oxidation state, of +3, fluorine in 2^{nd} period has close resemblances with chlorine of 3^{rd} 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 $_{19}K$ (having atomic number 19), the next element with similar properties is $_{37}Rb$.

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 $_{57}La$. 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 $_{87}Fr$ and $_{88}Ra$, seven transition elements and fourteen inner transition elements. The inner transition elements of this period are called **Actinides** as they follow $_{89}Ac$. 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 an 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.,

$$CaO + H_2O \longrightarrow Ca(OH)_2$$

$$Na_2O + H_2O \longrightarrow 2NaOH$$

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

$$M \longrightarrow M^+ + 1e^-$$

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.

$$N_2O_5 + H_2O \longrightarrow 2HNO_3$$

$$SO_3 + H_2O \longrightarrow H_2SO_4$$

$$P_2O_5 + 3H_2O \longrightarrow 2H_3PO_4$$

$$SO_2 + 2NaOH \longrightarrow Na_2SO_3 + H_2O$$

(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:

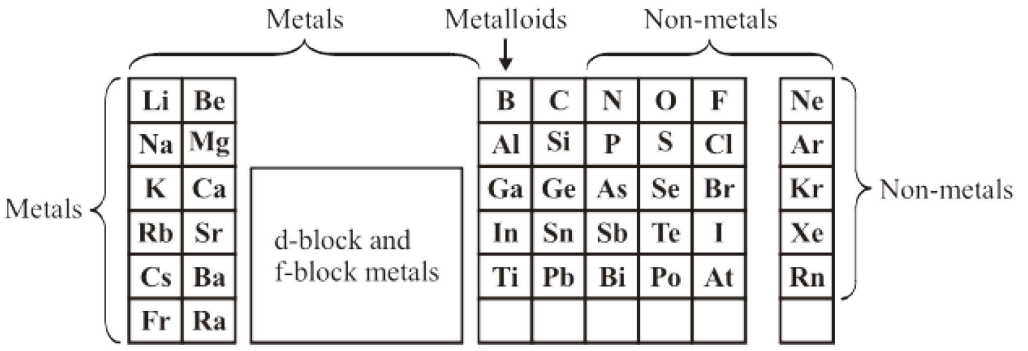
"The elements which have intermediate properties of metals and non-metals are called metalloids or semi-metals."

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.



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. Lowers 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.

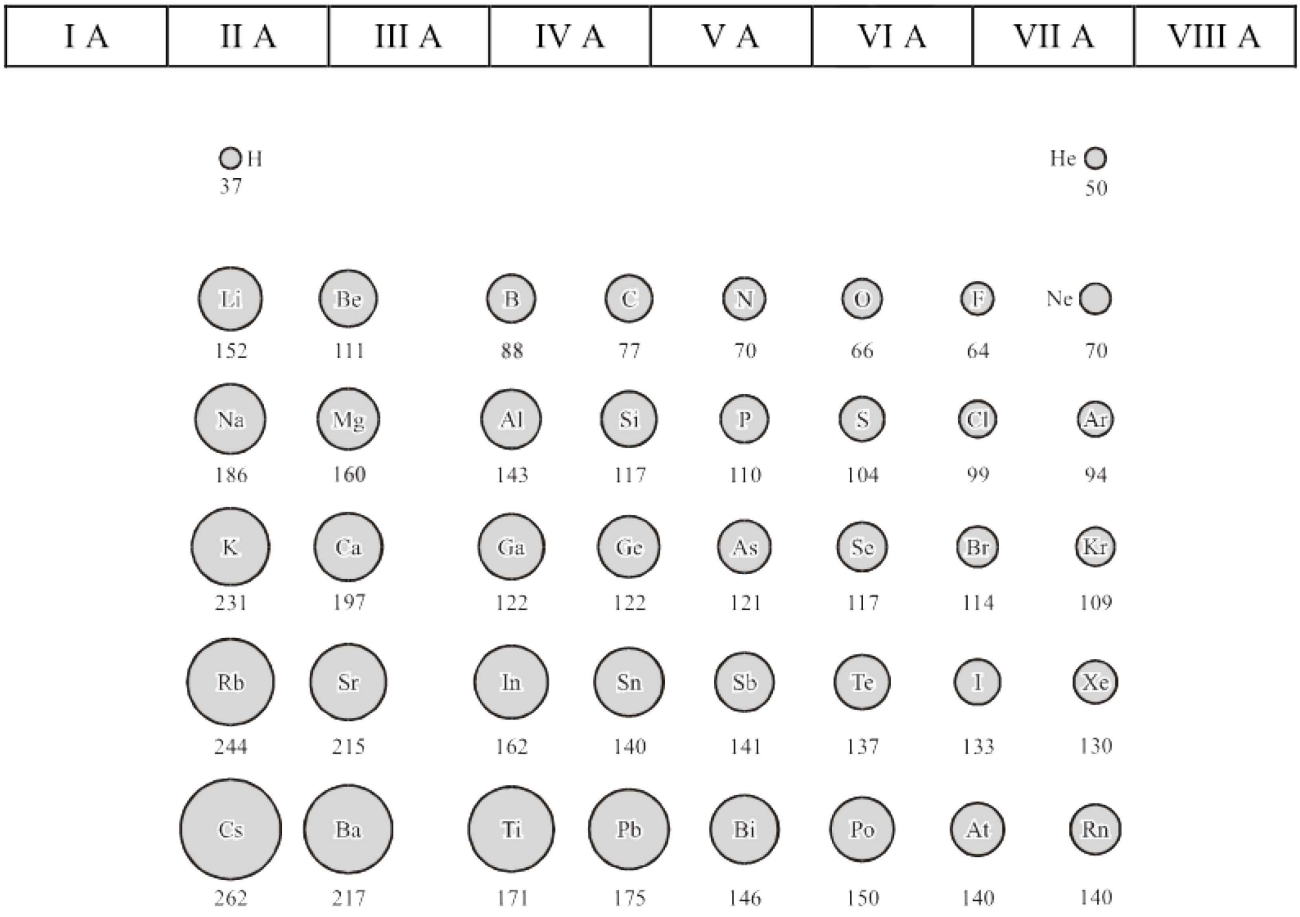
Size of a single atom cannot be measured directly because atoms have no sharp boundaries. Radius is defined as the distance form centre to boundary. We have studied that electrons have wave nature and concentration of cloud of electrons decreases rapidly

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.



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.

$$Na_{2,8,1} \longrightarrow Na^{+}_{2,8} + e^-$$

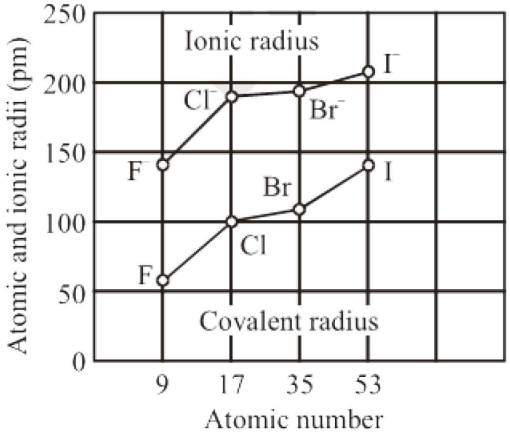
(2) By removal of electrons, the ratio of electrons to protons, do not remain balance. By removal of electron, the nuclear charge increases and remaining electrons are drawn closer to the nucleus. e.g., The covalent radius of Na is 157 pm and ionic radius of Na^+ is 95 pm (Atomic radius = 186 pm).

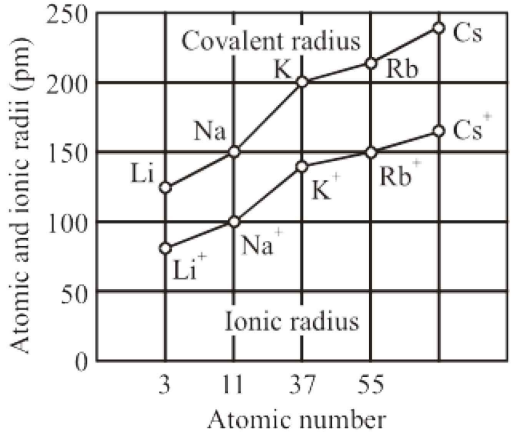
Size of Anions:

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 $\overline{F}$ 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}.

$$Na \longrightarrow Na^+ + e^- \qquad I_1 = 496 \text{ kJ mol}^{-1}$$

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

$$Mg \longrightarrow Mg^+ + 1e^- \qquad I_1 = 738 \text{ kJ mol}^{-1}$$

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

$$Mg^+ \longrightarrow Mg^{2+} + 1e^- \qquad I_2 = 1451 \text{ kJ mol}^{-1}$$

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.

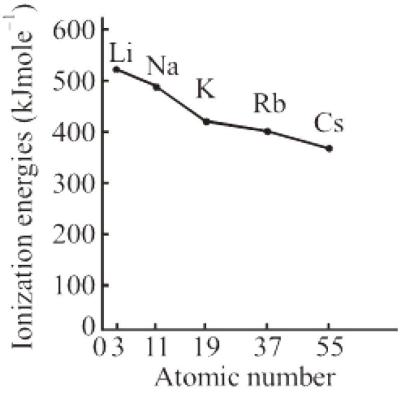
The shielding effect is actually the repulsion due to electrons in between the nucleus and the outermost shell. This effect increases, as the size of the atom increases due to addition of an extra shell successively in each period hence more number of electrons "shields" the nucleus. Going down in a 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-IE 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.



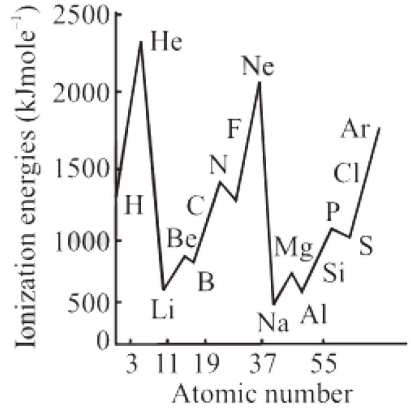


Table-I.E value of 2nd period (kJ per mole)

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.

$Be_4 = 1s^2, 2s^2$

$B_5 = 1s^2, 2s^2, 2p^1$

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

$N_{(7)} = 1s^2, 2s^2, 2px^1, 2py^1, 2pz^1$

$O_{(8)} = 1s^2, 2s^2, 2px^2, 2py^1, 2pz^1$

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.,

$$F + e^- \longrightarrow F^- \qquad E.A. = -337 \text{ kJ mol}^{-1}$$

$$Cl + e^- \longrightarrow Cl^- \qquad E.A = -349 \text{ kJ mol}^{-1}$$

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.

$$O + e^- \longrightarrow O^- \qquad E.A_1 = -141 \text{ kJ mol}^{-1}$$

$$O^- + e^- \longrightarrow O^{2-} \qquad E.A_2 = 780 \text{ kJ mol}^{-1}$$

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)	Ti −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

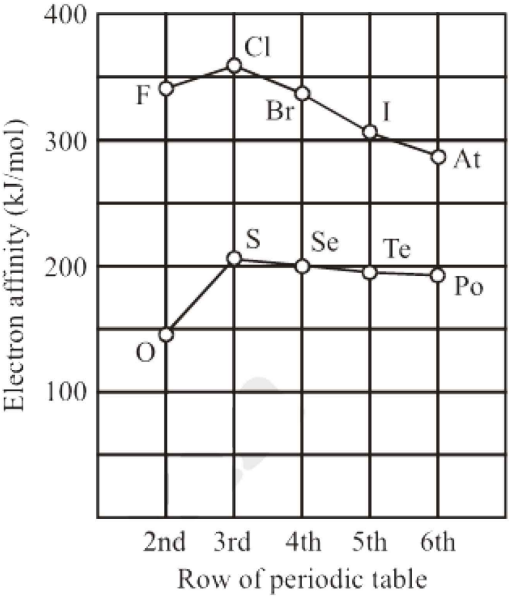
(a) Electron affinity depends upon:

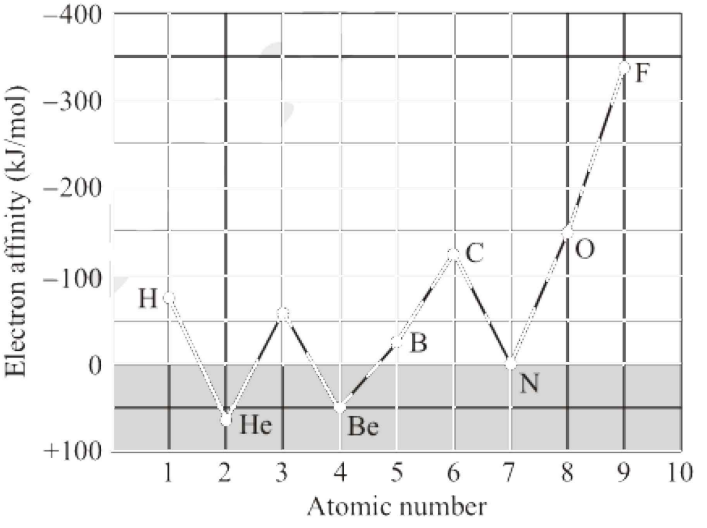
(i) Size of atoms

(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.,

$$Na_2O + H_2O \longrightarrow 2NaOH$$

$$K_2O + H_2O \longrightarrow 2KOH$$

$$CaO + H_2O \longrightarrow Ca(OH)_2$$

$$CaO + 2HCl \longrightarrow CaCl_2 + H_2O$$

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.

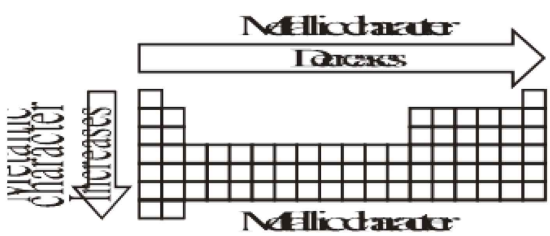
$$CO_2 + H_2O \longrightarrow H_2CO_3$$

$$SO_2 + H_2O \longrightarrow H_2SO_3$$

$$SO_3 + H_2O \longrightarrow H_2SO_4$$

$$SO_2 + 2NaOH \longrightarrow Na_2SO_3 + H_2O$$

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.

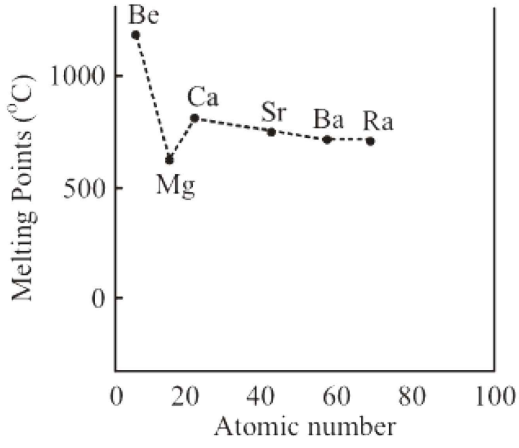
An important change occurs when we move from group IVA to VA, VIA, VIIA as the lighter elements of these groups exist as small, covalent molecules rather than as three dimensional lattices. For instance, nitrogen, oxygen and fluorine exist as individual diatomic molecules which have very weak intermolecular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

(b) Variation in Groups:

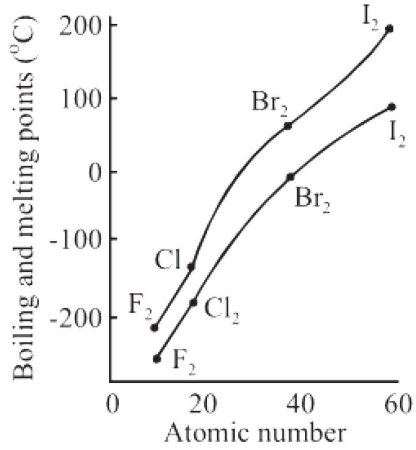
The melting and boiling points of **IA and IIA** group elements **decrease from top to bottom** due to the increase in their atomic sizes. The binding forces present between large sized atoms are relatively weaker as compared to those between smaller atoms.

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 oxdiation state of representaitve 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.,

$\overset{+1\ +5\ -6}{HNO_3}$, $\overset{+1\ +3\ -4}{HNO_2}$, $\overset{+3\ -3}{PCl_3}$, $\overset{+5\ -5}{PCl_5}$, $\overset{+3\ +5\ -8}{H_3PO_4}$, $\overset{+3\ +5\ -6}{H_3PO_3}$

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

$\overset{+4\ -4}{SO_2}$, $\overset{+6\ -6}{SO_3}$, $\overset{+2\ +4\ -6}{H_2SO_3}$, $\overset{+2\ +6\ -8}{H_2SO_4}$

(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.,

$\overset{+1\ +7\ -8}{HClO_4}$, $\overset{+1\ +5\ -6}{HClO_3}$, $\overset{+1\ +3\ -4}{HClO_2}$, $\overset{+1\ +1\ -2}{HClO}$

(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:

$\overset{+2\ -2}{Cu_2Cl_2}$, $\overset{+2\ +6\ -8}{ZnSO_4}$, $\overset{+10\ -10}{V_2O_5}$, $\overset{+2\ +6\ -8}{K_2CrO_4}$, $\overset{+1\ +7\ -8}{KMnO_4}$, $\overset{+2\ +6\ -8}{K_2MnO_4}$

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 tetrahydrally 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 on 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.

$$H^{+}_{(g)} + \text{Water} \longrightarrow H^{+}_{(aq)} \qquad \Delta H_h = -1075 \text{ kJ mol}^{-1}$$

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

Table – Hydration energies of ions

Ion	ΔH_h kJ mol^{-1}	Ion	ΔH_h kJ mol^{-1}	Ion	ΔH_h kJ mol^{-1}	Ion	ΔH_h 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:

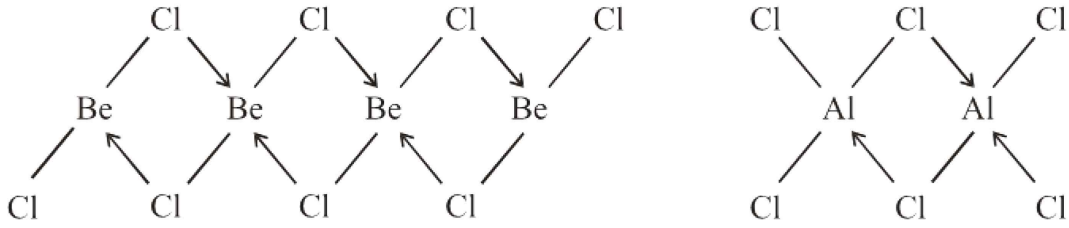
"If ionic bond is present between the metal and halogen, the halides are called ionic halides."

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 > Bormide > 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 in 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_2Cl_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 fairy 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 hybrides."

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 $\bar{H}$ (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_2O	HF
NaH	MgH_2		AlH_3	SiH_4	PH_3	H_2S	HCl
KH	CaH_2	ZnH_2	GaH_3	GeH_4	AsH_3	H_2Se	HBr
RbH	SrH_2	CdH_2	InH_3	SnH_4	SbH_3	H_2Te	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 hybride and the least stable are those of thallium, lead and bismuth.

(5) These hybrides are formed by elements with electronegativity value greater than 1.8 (Pauling scale). Since the electronegativity of hydrogen is 2.1, most of these hybrides 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. (°C)	–184	–185	–165	–150
B.P. (°C)	–164	–112	–90	–52

Hydrides (Group VIA)	**H_2O**	**H_2S**	**H_2Se**	**H_2Te**
M.P. (°C)	0.00	–82.9	–65.7	–48
B.P. (°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:

$$Na_2O + CO_2 \longrightarrow Na_2CO_3$$

$$Na_2O + SO_3 \longrightarrow Na_2SO_4$$

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

Basic Oxides:

Oxides of **IA and IIA group** elements from **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 changes to hydroxide ions OH^-.

$$Na_2O + H_2O \longrightarrow 2NaOH$$

$$CaO + H_2O \longrightarrow Ca(OH)_2$$

Amphoteric Oxides:

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

(i)

$$ZnO + H_2SO_4 \longrightarrow ZnSO_4 + H_2O$$

$$ZnO + 2NaOH + H_2O \longrightarrow Na_2[Zn(OH)_4]$$

(ii)

$$Al_2O_3 + 6HCl \longrightarrow 2AlCl_3 + 3H_2O$$

$$Al_2O_3 + 2NaOH \longrightarrow 2NaAlO_2 + H_2O$$

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.,

$$CO_2 + H_2O \longrightarrow H_2CO_3$$

$$SO_3 + H_2O \longrightarrow H_2SO_4$$

$$N_2O_5 + H_2O \longrightarrow 2HNO_3$$

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:

$\underbrace{Na_2O, MgO}_{\text{Basic}}$ $\quad Al_2O_3$ $\quad \underbrace{SiO_2, P_4O_{10}, SO_3, Cl_2O_7}_{\text{Acidic}}$

Amphoteric

The basic characters of IA and IIA metal oxides increase downward in a group BeO < MgO < SrO < 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 MnO < Mn_2O_3 < MnO_2 < Mn_2O_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	$(M^+)_2O^{2-}$	$M^{+2}O^{-2}$	$M^{+2}O^{-2}$	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	TI	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.,

$$Li = 1s^2, 2s^1$$

$$Na = 1s^2, 2s^2, 2p^6, 3s^1$$

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

$$H \longrightarrow H^+ + e^-$$

$$Na \longrightarrow Na^+ + e^-$$

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

$$H_2 + Cl_2 \longrightarrow 2HCl$$

$$2Na + Cl_2 \longrightarrow 2NaCl$$

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

$$Na^+H^-, Ca^{+2}H_2^{-2}$$

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^+.

$$H^+ + H_2O \longrightarrow H_3O^+$$

(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.

$F_{(g)} = 1s^2, 2s^2, 2p^5$

$Cl_{(17)} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^5$

(2) Hydrogen is a gas like most of the halogens.

(3) Hydrogen and halogens exist in diatomic state. e.g.,

H_2, F_2, Cl_2, Br_2 and I_2

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

$H + 1e^- \longrightarrow H^-$ (Hydride ion)

$Cl + 1e^- \longrightarrow Cl^-$ (Chloride ion)

(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.

$H^- + H_2O \longrightarrow OH^- + H_2$

(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.

$C_6 = 1s^2, 2s^2, p^2$

$S_{14} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^2$

(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.

$CuO + H_2 \longrightarrow Cu + H_2O$

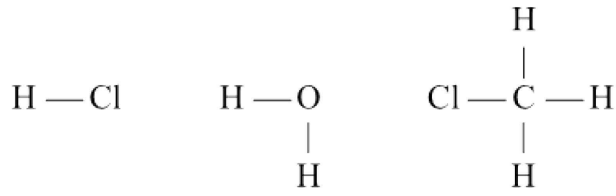
$SnO_2 + C \longrightarrow Sn + CO_2$

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.

EXERCISE

Q.1 Fill in the blanks:

(i) Mendeleev in his periodic table, arranged the elements according to their atomic ————.

(ii) Vertical columns in modern periodic table are called ———— and horizontal rows are called ————.

(iii) Members of group VIIA are called ———— and alkali metals is the family name of group member ————.

(iv) Metals form ———— oxides and non-metals form ———— oxides.

(v) Hydrogen can be placed above the groups ———— of the periodic table.

(vi) Shielding effect is actually the ———— due to electrons in between the nucleus and the outermost shell.

(vii) Noble gases have the ———— values of ionization energy due to their complete outermost shells.

(viii) When a second electron is added to a uninegative ion, the incoming electron is ———— by the already present negative charge.

(ix) Due to having partly filled d-orbitals ———— metals usually show variable valency.

(x) Melting and boiling points of halogens ———— down the group.

ANSWERS

(i)	masses	**(ii)**	group, period	**(iii)**	Halogens, IA
(iv)	basic, acidic	**(v)**	IA, IVA, VIIA	**(vi)**	repulsion
(vii)	high	**(viii)**	repelled	**(ix)**	transition
(x)	increase				

Q.2 Indicate True or False:

(i) In Mendeleev's periodic table elements Be, Mg, Zn and Cd are placed in the same group.

(ii) The second and third periods contain eighteen elements each.

(iii) Alkaline earth metals are present in Group IIA.

(iv) Metals are present in the top right corner of the periodic table.

(v) Metalloids are present in the lower half of Groups IVA, VA and VIA.

(vi) Hydrogen forms uninegative ion like halogens.

(vii) Oxidation state of an element is related to the number of period it belongs.

(viii) Diamond is a good conductor of electricity.

(ix) Melting points of halogens decrease down the group.

(x) Zinc oxide is an example of amphoteric oxide.

ANSWERS

(i)	True	**(ii)**	False	**(iii)**	True	**(iv)**	False	**(v)**	True
(vi)	True	**(vii)**	False	**(viii)**	False	**(ix)**	False	**(x)**	True

Q.3 Multiple choice questions. Encircle the correct answer:

(i) Keeping in view the size of atoms, which orders is the correct one:

(a) Mg > Sr **(b)** Ba > Mg

(c) Lu > Ce **(d)** Cl > I

(ii) Mark the correct statement:

(a) Na^+ is smaller than Na atom **(b)** Na^+ is larger than Na atom

(c) Cl^- is smaller than Cl atom **(d)** Cl^- (ion) and Cl (atom) are equal in size

(iii) Mark the correct statement:

(a) All lanthanides are present in the same group

(b) All halogens are present in the same period

(c) All the alkali metals are present in the same group

(d) All the noble gases are present in the same period

(iv) Which statement is incorrect:

(a) All the metals are good conductor of electricity

(b) All the metals are good conductor of heat

(c) All the metals form positive ions

(d) All the metals form acidic oxides

(v) Which statement is correct:

(a) Hydrogen resembles in properties with IA, IVA and VIIA elements

(b) Hydrogen resembles in properties with IIIA, IVA and VA elements

(c) Hydrogen resembles in properties with IIA, IVA and VIA elements

(d) Hydrogen resembles in properties with IIA, IIIA and VIIA elements

(vi) Mark the correct statement:

(a) The ionization energy of calcium is lower than that of barium

(b) The ionization energy of calcium is lower than that of magnesium

(c) The ionization energy of calcium is higher than that of beryllium

(d) The ionization energy of calcium is lower than that of strontium

(vii) Mark the correct statement:

(a) Electron affinity is a measure of energy required to remove the electron

(b) Electron affinity is a measure of energy released by adding an electron

(c) Electron affinity is a measure of energy required to excite an electron

(d) Electron affinity is measure of energy released by removing an electron

(viii) Mark the correct statement:

(a) Metallic character increases down the group

(b) Metallic character increases from left to right along a period

(c) Metallic character remains the same from left to right along a period

(d) Metallic character remains the same down the group

(ix) Mark the correct statement:

(a) Melting points of halogens decrease down the group

(b) Melting points of halogens increase down the group

(c) Melting points of halogens remain the same throughout the group

(d) Melting points of halogens first increase and then decrease down he group

(x) Mark the correct statement:

(a) Covalent character of metal halides increases from left to right in a period

(b) Boiling points of group IVA hydrides decrease down the group

(c) Ionic character of hydrides increases from left to right in a period

(d) The basicity of group IIA oxides decreases on descending the group

ANSWERS

(i)	(b)	**(ii)**	(a)	**(iii)**	(c)	**(iv)**	(d)	**(v)**	(a)
(vi)	(b)	**(vii)**	(b)	**(viii)**	(a)	**(ix)**	(b)	**(x)**	(a)

Q.4 What are the improvements made in the Mendeleev's periodic table?

Ans. The improvements in Mendeleev's periodic table are as followings:

(i) The periodic table was improved by arranging the elements in ascending order of their atomic numbers instead of their atomic masses.

(ii) The addition of an extra group (VIII-A) at the extreme right of the periodic table. This group contains noble gases which had not been discovered in Mendeleev's time.

(iii) In modern periodic table the elements of single vertical group of Mendeleev's periodic table is divided into two types of vertical groups, A and B. e.g., In Mendeleev's periodic table Be, Mg, Ca, Sr, Ba and Zn, Cd, Hg were placed in a single group but in modern periodic table Be, Mg, Ca, Sr and Ba are placed in group II-A and Zn, Cd, Hg in II-B group.

Q.5 How the classification of elements in different blocks helps in understanding their chemistry?

Ans. This classification is based upon the valence orbital of the element involved in chemical bonding. So this classification is quite useful in understanding the chemical properties of elements especially the concept of valency as oxidation state.

Q.6 How do you justify the position of hydrogen at the top of various groups?

Ans. Descriptive question. See text book for details.

Q.7 Why the ionic radii of negative ions are larger than the size of their parent atoms?

Ans. See Q.1 of Short Question.

Q.8 Why ionization energy decreases down the group and increases along a period?

Ans. As we know that:

$$\text{Ionization energy} \propto \text{Nuclear force} \propto \frac{1}{\text{Atomic size}}$$

(i) As we go down the group the atomic size increases which results as the decrease in nuclear force due to which ionization energy decreases down the group.

(ii) While moving from left to right in a period, the nuclear force increases and atomic size decreases which makes the removal of electron difficult and hence the value of ionization energy increases.

Q.9 Why the second value of electron affinity of an element is usually shown with a positive sign?

Ans. When a second electron is added to a uni-negative ion, the incoming electron is repelled by the already present negative charge and energy is absorbed in this process so, it is usually shown with a positive sign.

Q.10 Why metallic character increases from top to bottom in a group of metals?

Ans. Metallic character increases from top to bottom in groups due to increase in size of atom and ease to remove electrons having low ionization energy.

Q.11 Explain the variation in melting points along the short periods.

Ans. Melting points in short periods increase upto the middle and than decreases.

(i) Increase in melting point is due to decrease in atomic size and increase in inter-atomic forces.

(ii) After group IV-A lighter elements of these groups exist as small, covalent molecules and have weak inter-molecular forces due to which melting points also decreases.

Q.12 Why the oxidation state of noble gases is usually zero?

Ans. Noble gases consist of He, Ne, Ar, Kr, Xe and Rn. Their atomic numbers are 2, 10, 18, 36, 54 and 86 respectively. As the oxidation state of an element usually depends on the number of electrons which it contributes in bond formation. Noble gases have completely filled outermost shell and they have zero oxidation state.

Q.13 Why diamond is a non-conductor and graphite is fairly a good conductor?

Ans. Electrical conductance depends upon the number of free electrons. In case of diamond all the valence electrons are used for covalent bond formation and no free electrons are available for conduction of electricity and hence diamond is a non-conductor. In case of graphite carbon is fairly good conductor because one of its four valence electrons is relatively free to move and hence can conduct electricity.

Q.14 Give brief reason for the following:

(a) d and f-block elements are called transition elements?

(b) Lanthanide contraction controls the atomic sizes of elements of 6^{th} and 7^{th} periods.

(c) The melting and boiling points of the elements increase from left to right up to the middle of p-block elements and decrease onward.

(d) The oxidation states vary in a period but remain almost constant in a group.

(e) The hydration energies of the ions are in the following order:

$$Al^{3+} > Mg^{2+} > Na^{+}$$

(f) Ionic character of halides decreases from left to the right in a period.

(g) Alkali metals give ionic hydrides.

(h) Although both sodium and phosphorus are present in the same period of the periodic table yet their oxides are different in nature, Na_2O is basic while P_2O_5 is acidic in character.

Ans.

(a) d and f-block elements are less electropositive than alkali and alkaline earth metals. They have intermediate characters between s and p-block elements and hence are called transition elements.

(b) In case of Lanthanide and Actinide series there is a gradual decrease in the atomic size from left to right due to increase in nuclear charge. This gradual decrease in atomic size in these series is known as **lanthanide contraction** and it controls the size or 6^{th} and 7^{th} period elements.

(c) Elements of I-A to IV-A group exist as three dimensional, giant structures and number of valence electrons increases up to group IV-A due to which melting points and boiling points increases up to IV-A group. On the other hand elements of V-A up to VII- A group exist as non-polar, small and diatomic molecules having weak intermolecular forces fore melting and boiling points of V-A to onward elements decreases from left to right.

(d) As the oxidation states of elements is based upon the electronic configuration of elements. In a period of periodic table, the electronic configuration changes but in a group the electronic configuration remains constant. Therefore due to same electronic configuration in elements of same group, they have same oxidation states.

(e) Hydration energy of an ion depends upon the **charge** and the **size** of ion. It is directly proportional to ionic charge and inversely proportional to its size. Al^{3+}, Mg^{2+} and Na^{+} ions are iso-electronic and their size decreases in the following order:

$$Na^{+} > Mg^{2+} > Al^{3+}$$

So hydration energy of Al^{3+} will be maximum among these ions. Hydration energies of these ions are given below:

Na^{+} = –390 KJ/mol

Mg^{2+} = –1891 KJ/mol

Al^{3+} = –4613 KJ/mol

(f) Ionic character of halides decreases along a period due to decrease in difference of electro negativity. e.g., In 3rd period:

(i) NaCl and $MgCl_2$ are ionic halide.

(ii) $AlCl_3$ is polymeric halide.

(iii) $SiCl_4$, PCl_3, S_2Cl_2 are covalent in nature.

(g) Alkali metals are electropositive elements and loose electron easily.

$$M \longrightarrow M^{+} + e^{-}$$

These metals have low ionization energies. Hydrogen atom accepts one electron and form hydride ion.

$$H + e^{-} \longrightarrow H^{-}$$

These metal ions and hydride ion combine to form ionic hydrides. e.g.,

$$2Na + H_2 \longrightarrow 2NaH$$

(h) Na is a metal while P is a non-metal. Metals form basic oxides while non-metals form acidic oxides. Na_2O when dissolved in water form NaOH

$$Na_2O + H_2O \longrightarrow 2NaOH(base)$$

Base oxide

while P_2O_5 when dissolved in water form H_3PO_4.

$$P_2O_5 + 3H_2O \longrightarrow 2H_3PO_4 \quad (acid)$$

Acidic oxide

DO YOU KNOW?

(1) Al-Razi classified the elements on the basis of physical and chemical properties of the substances.

(2) Dobrenier, a German chemist in 1929 gave the law of triads.

(3) Mendleev corrected the atomic mass of Be and placed it in the II group above Mg, Ca, Sr, Ba.

(4) The elements from Ce_{58} — Lu_{71} are called Lanthnides.

(5) The elements from Th_{90} — Lr_{103} are called actinides.

(6) Cu, Ag, Au of group I-B are called coinage metals.

(7) The elements after U_{92} are called transuranic elements.

(8) The elements of group VII are called halogens.

(9) The elements of group VI are called chalcogens (ore forming).

(10) Oxidation state of O in OF_2 is +2.

(11) AlF_3 is ionic compound (M.P = 1290°C) but AlI_3 is covalent in nature (M.P = 198°C).

(12) $PbCl_2$ is ionic while $PbCl_4$ is covalent.

(13) Ionization energy of B (801 kJ mol^{-1}) is less than Be (899 kJ mol^{-1}).

(14) Ionization energy of O (1314 kJ mol^{-1}) is less than N (1402 kJ mol^{-1}).

(15) Ionization energy of K (419) is less than Ca (590) (Units of I.E = kJ/m)

(16) The oxides of Be, Zn, Al, Pb are amphoteric.

(17) CH_4 is neutral hydride.

(18) NH_3 is basic hydride.

(19) Cesium is most reactive metal.

(20) Fluorine is most reactive non-metal.

(21) Highest oxidation state (Mn^{+7}).

(22) Helium has highest ionization energy.

(23) Cesium is strong reducing agent.

(24) Heaviest solid is osmium.

(25) Lightest gas is hydrogen.

(26) Lanthanides and actinides are called rare earth elements.

CHAPTER 1

PERIODIC CLASSIFICATION OF ELEMENTS AND PERIODICITY

MULTIPLE CHOICE QUESTIONS

1. 7^{th} period of the periodic table contains normal elements:

(a) 2 **(b)** 3

(c) 4 **(d)** 8

2. Periodic table has been divided into four blocks, which blocks contain highest elements:

(a) s **(b)** p

(c) d **(d)** f

3. The division of elements into blocks in the modern periodic table is based on:

(a) Shell **(b)** Sub-shell

(c) Orbital **(d)** All

4. Keeping in view the size of atoms, which order is the correct one:

(a) Mg > Sr **(b)** Ba > Mg

(c) Lu > Ce **(d)** Cl > I

5. Mosley concluded by the X-ray studies of the elements that physical and chemical properties of elements depend upon:

(a) Atomic mass **(b)** Atomic no.

(c) Mass no. **(d)** All

6. Which one of the following is not ionic hydride:

(a) CsH **(b)** LiH

(c) HCl **(d)** NaH

7. The element with highest first ionization energy is:

(a) B **(b)** C

(c) O **(d)** N

8. Mark the correct statement:

(a) Na^+ is smaller than Na **(b)** Na^+ is larger than Na

(c) Cl^- is smaller than Cl **(d)** Cl^- ion and Cl are equal

9. The element with atomic number 9 is closest in chemical properties with element of atomic-number:

(a) 27 **(b)** 11

(c) 17 **(d)** 8

10. Period six in periodic table contain which "block" elements:

(a) s, p **(b)** s, p, d

(c) s, p, f **(d)** s, p, d, f

11. Who introduced the zero group:

(a) Lother Mayer **(b)** Mendeleev

(c) Ramsay **(d)** Mosley

12. Elements of group II-B are called:

(a) Representative elements **(b)** Transition elements

(c) Non-typical transition elements **(d)** Coinage metal group

13. The elements with lowest M. Pt:

(a) Be **(b)** Mg

(c) Cd **(d)** Cr

14. Which of the following sets of atoms are arranged in order of decreasing electronegativity?

(a) N, O, F **(b)** Si, P, S

(c) F, O, N **(d)** S, Si, P

15. When hydrogen loses its electron to form H^+ ion then it resemble:

(a) Transition metals **(b)** Halogen

(c) Alkali metals **(d)** Noble gases

16. Which of the halides show bridge type structure:

(a) NaCl **(b)** $AlCl_3$

(c) CCl_4 **(d)** H_2O

17. Which element of the following shows max. oxidation state:

(a) P **(b)** S

(c) Mn **(d)** Cr

18. M.Pt is highest for the elements of group:

(a) IA **(b)** IIA

(c) IIIA **(d)** IVA

19. Amphoteric oxide is formed by:

(a) Ca **(b)** Fe

(c) Cu **(d)** Zn

20. In the long form of the periodic table, the most electropositive elements occupy:

(a) Bottom right position **(b)** Bottom left position

(c) Top right position **(d)** Top left position

answers

1.	(a)	**2.**	(c)	**3.**	(b)	**4.**	(b)	**5.**	(b)
6.	(c)	**7.**	(d)	**8.**	(a)	**9.**	(c)	**10.**	(d)
11.	(d)	**12.**	(c)	**13.**	(b)	**14.**	(c)	**15.**	(c)
16.	(b)	**17.**	(c)	**18.**	(d)	**19.**	(d)	**20.**	(b)

SOLVED EXERCISE

Q.1 Fill in the blanks:

(i) Mendeleev in his periodic table, arranged the elements according to their atomic ————.

(ii) Vertical columns in modern periodic table are called ———— and horizontal rows are called ————.

(iii) Members of group VIIA are called ———— and alkali metals is the family name of group member ————.

(iv) Metals form ———— oxides andnon-metals form ———— oxides.

(v) Hydrogen can be placed above the groups ———— of the periodic table.

(vi) Shielding effect is actually the ———— due to electrons in between the nucleus and the outermost shell.

(vii) Noble gases have the ———— values of ionization energy due to their complete outermost shells.

(viii) When a second electron is added to a uninegative ion, the incoming electron is ———— by the already present negative charge.

(ix) Due to having partly filled d-orbitals ———— metals usually show variable valency.

(x) Melting and boiling points of halogens ———— down the group.

answers

(i)	masses	**(ii)**	group, period	**(iii)**	Halogens, IA
(iv)	basic, acidic	**(v)**	IA, IVA, VIIA	**(vi)**	repulsion
(vii)	high	**(viii)**	repelled	**(ix)**	transition
(x)	increase				

Q.2 Indicate True or False:

(i) In Mendeleev's periodic table elements Be, Mg, Zn and Cd are placed in the same group.

(ii) The second and third periods contain eighteen elements each.

(iii) Alkaline earth metals are present in Group IIA.

(iv) Metals are present in the top right corner of the periodic table.

(v) Metalloids are present in the lower half of Groups IVA, VA and VIA.

(vi) Hydrogen forms uninegative ion like halogens.

(vii) Oxidation state of an element is related to the number of period it belongs.

(viii) Diamond is a good conductor of electricity.

(ix) Melting points of halogens decrease down the group.

(x) Zinc oxide is an example of amphoteric oxide.

answers

(i)	True	**(ii)**	False	**(iii)**	True	**(iv)**	False	**(v)**	True
(vi)	True	**(vii)**	False	**(viii)**	False	**(ix)**	False	**(x)**	True

Q.3 Multiple choice questions. Encircle the correct answer:

(i) Keeping in view the size of atoms, which orders is the correct one:

(a) Mg > Sr **(b)** Ba > Mg

(c) Lu > Ce **(d)** Cl > I

(ii) Mark the correct statement:

(a) Na^+ is smaller than Na atom **(b)** Na^+ is larger than Na atom

(c) Cl^- is smaller than Cl atom **(d)** Cl^- (ion) and Cl (atom) are equal in size

(iii) Mark the correct statement:

(a) All lanthanides are present in the same group

(b) All halogens are present in the same period

(c) All the alkali metals are present in the same group

(d) All the noble gases are present in the same period

(iv) Which statement is incorrect:

(a) All the metals are good conductor of electricity

(b) All the metals are good conductor of heat

(c) All the metals form positive ions

(d) All the metals form acidic oxides

(v) **Which statement is correct:**

(a) Hydrogen resembles in properties with IA, IVA and VIIA elements

(b) Hydrogen resembles in properties with IIIA, IVA and VA elements

(c) Hydrogen resembles in properties with IIA, IVA and VIA elements

(d) Hydrogen resembles in properties with IIA, IIIA and VIIA elements

(vi) **Mark the correct statement:**

(a) The ionization energy of calcium is lower than that of barium

(b) The ionization energy of calcium is lower than that of magnesium

(c) The ionization energy of calcium is higher than that of beryllium

(d) The ionization energy of calcium is lower than that of strontium

(vii) **Mark the correct statement:**

(a) Electron affinity is a measure of energy required to remove the electron

(b) Electron affinity is a measure of energy released by adding an electron

(c) Electron affinity is a measure of energy required to excite an electron

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answers

(i)	(b)	**(ii)**	(a)	**(iii)**	(c)	**(iv)**	(d)	**(v)**	(a)
(vi)	(b)	**(vii)**	(b)	**(viii)**	(a)	**(ix)**	(b)	**(x)**	(a)

Q.4 What are the improvements made in the Mendeleev's periodic table?

Ans. The improvements in Mendeleev's periodic table are as followings:

(i) The periodic table was improved by arranging the elements in ascending order of their atomic numbers instead of their atomic masses.

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Ans. This classification is based upon the valence orbital of the element involved in chemical bonding. So this classification is quite useful in understanding the chemical properties of elements especially the concept of valency as oxidation state.

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Ans. Descriptive question. See text book for details.

Q.7 Why the ionic radii of negative ions are larger than the size of their parent atoms?

Ans. See Q.1 of Short Question.

Q.8 Why ionization energy decreases down the group and increases along a period?

Ans. As we know that:

$$\text{Ionization energy} \propto \text{Nuclear charge} \propto \frac{1}{\text{Atomic size}}$$

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Ans. Melting points in short periods increase upto the middle and than decreases.

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(d) The oxidation states vary in a period but remain almost constant in a group.

(e) The hydration energies of the ions are in the following order:

$$Al^{3+} > Mg^{2+} > Na^{+}$$

(f) Ionic character of halides decreases from left to the right in a period.

(g) Alkali metals give ionic hydrides.

(h) Although both sodium and phosphorus are present in the same period of the periodic table yet their oxides are different in nature, Na_2O is basic while P_2O_5 is acidic in character.

Ans.

(a) d and f-block elements are less electropositive than alkali and alkaline earth metals. They have intermediate characters between s and p-block elements and hence are called transition elements.

(b) In case of Lanthanide and Actinide series there is a gradual decrease in the atomic size from left to right due to increase in nuclear charge. This gradual decrease in atomic size in these series is known as **lanthanide contraction** and it controls the size or 6^{th} and 7^{th} period elements.

(c) Elements of I-A to IV-A group exist as three dimensional, giant structures and number of valence electrons increases up to group IV-A due to which melting points and boiling points increases up to IV-A group. On the other hand elements of V-A up to VII- A group exist as non-polar, small and diatomic molecules having weak intermolecular forces for melting and boiling points of V-A to onward elements decreases from left to right.

(d) As the oxidation states of elements is based upon the electronic configuration of elements. In a period of periodic table, the electronic configuration changes but in a group the electronic configuration remains constant. Therefore due to same electronic configuration in elements of same group, they have some oxidation states.

(e) Hydration energy of an ion depends upon the **charge** and the **size** of ion. It is directly proportional to ionic charge and inversely proportional to its size. Al^{3+}, Mg^{2+} and Na^{+} ions are iso-electronic and their size decreases in the following order:

$$Na^{+} > Mg^{2+} > Al^{3+}$$

So hydration energy of Al^{3+} will be maximum among these ions. Hydration energies of these ions are given below:

Na^{+} = –390 KJ/mol

Mg^{2+} = –1891 KJ/mol

Al^{3+} = –4613 KJ/mol

(f) Ionic character of halides decreases along a period due to decrease in difference of electro negativity. e.g., In 3rd period:

(i) NaCl and $MgCl_2$ are ionic halide.

(ii) $AlCl_3$ is polymeric halide.

(iii) $SiCl_4$, PCl_3, S_2Cl_2 are covalent in nature.

(g) Alkali metals are electropositive elements and loose electron easily.

$$M \longrightarrow M^{+} + e^{-}$$

These metals have low ionization energies. Hydrogen atom accepts one electron and form hydride ion.

$$H + e^{-} \longrightarrow H^{-}$$

These metal ions and hydride ion combine to form ionic hydrides. e.g.,

$$2Na + H_2 \longrightarrow 2NaH$$

(h) Na is a metal while P is a non-metal. Metals form basic oxides while non-metals form acidic oxides. Na_2O when dissolved in water form NaOH

$$Na_2O + H_2O \longrightarrow 2NaOH \quad \text{(base)}$$

Base oxide

while P_2O_5 when dissolved in water form H_3PO_4.

$$P_2O_5 + 3H_2O \longrightarrow 2H_3PO_4 \quad \text{(acid)}$$

Acidic oxide

position of hydrogen in the periodic table

Hydrogen is not a metal but in most of versions of 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 outer most shell and its electronic configuration is $1s^1$. Alkali metals also have electronic configuration ns^1 e.g.,

   $Li = 1s^2, 2s^1,\ Na = 1s^2, 2s^2, 2p^6, 3s^1$

2. Hydrogen loses one electron like alkali metals and form H^+ ion:

   $H \longrightarrow H^+ + e^-$

   $Na \longrightarrow Na^+ + e^-$

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

   $H_2 + Cl_2 \longrightarrow 2HCl$

   $2Na + Cl_2 \longrightarrow 2NaCl$

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

   $Na^+H^-,\ Ca^{+2}H_2^{-1(2)}$

(b) 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 metal ions, e.g., Na^+, K^+ etc. H^+ associates with H_2O to form H_3O^+.

   $H^+ + H_2O \longrightarrow H_3O^+$

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.

(c) 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.

$F_{(9)} = 1s^2, 2s^2, 2p^5$

$Cl_{(17)} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^5$

2. Hydrogen is a gas like most of the halogens.

3. Hydrogen and halogens exist in diatomic state, e.g., H_2, F_2, Cl_2, Br_2 and I_2.

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

$H + 1e^- \longrightarrow H^-$ (Hydride ion)

$Cl + 1e^- \longrightarrow Cl^-$ (Chloride ion)

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

(d) Dissimilarities of hydrogen 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 combines with oxygen to form very stable oxide (H_2O) while halogens does not form such stable oxides.

4. Halide ions (F^-, Cl^-, Br^-, I^-) are stable in aqueous form while H^- is unstable in water.

$H^- + H_2O \longrightarrow OH^- + H_2$

(e) 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.

$C_6 = 1s^2, 2s^2, 2p^2$

$S_{14} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^2$

2. Hydrogen and carbon combine with other elements with covalent bond.

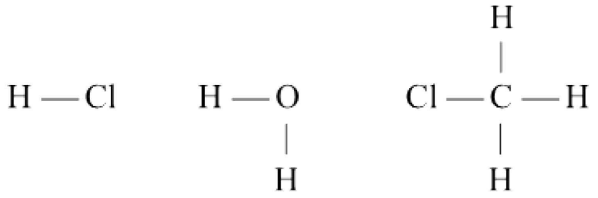
3. In majorities of reactions of carbon and hydrogen, they act as reducing agents,

$CuO + H_2 \longrightarrow Cu + H_2O$

$SnO_2 + C \longrightarrow Sn + CO_2$

(f) Dissimilarities of hydrogen with group IVA:

1. Carbon and silicon form chain type compounds with each other or with other elements 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.



4. Hydrogen is a gas while group IV elements are all solids.
5. Hydrogen is a s-block while carbon is a p-block element.

DO YOU KNOW?

1. Al-Razi classified the elements on the basis of physical and chemical properties of the substances.
2. Dobrenier, a German chemist in 1929 gave the law of triads.
3. Mendleev corrected the atomic mass of Be and placed it in the II group above Mg, Ca, Sr, Ba.
4. The elements from Ce_{58} — Lu_{71} are called Lanthnides.
5. The elements from Th_{90} — Lr_{103} are called actinides.
6. Cu, Ag, Au of group I-B are called coinage metals.
7. The elements after U_{92} are called transuranic elements.
8. The elements of group VII are called halogens.
9. The elements of group VI are called chalcogens (ore forming).
10. Oxidation state of O in OF_2 is +2.
11. AlF_3 is ionic compound (M.P = 1290°C) but AlI_3 is covalent in nature (M.P = 198°C).
12. $PbCl_2$ is ionic while $PbCl_4$ is covalent.
13. Ionization energy of B (801 kJ mol^{-1}) is less than Be (899 kJ mol^{-1}).
14. Ionization energy of O (1314 kJ mol^{-1}) is less than N (1402 kJ mol^{-1}).
15. Ionization energy of K (419) is less than Ca (590) (Units of I.E = KJ/m)
16. The oxides of Be, Zn, Al, Pb are amphoteric.

SHORT QUESTIONS

Q.1 Why ionic radii of anions are larger than their parent atoms?

Ans. There are the following reasons:

(a) Incoming electron is accommodated in the new shell thus increasing the orbit and therefore size increases.

(b) There is decrease in the nuclear charge.

(c) If the incoming electron enters the outermost orbit which has already electrons, the orbit will expand. For example atomic radius of F is 72 Pm while ionic radius of F^- is 136 Pm.

Q.2 Why ionization energy decreases down the group?

Ans. It is due to:

(a) There is increase in the atomic size down the group, the hold of the nucleus on the electrons decreases.

(b) Shielding effect also increases.

Q.3 Why ionization energy increases along the period?

Ans. (a) Along the period, there is decrease in the atomic size, hold of the nucleus on the electrons increases.

(b) Nuclear charge increases.

Q.4 Why d and f block elements are called transition elements?

Ans. The elements of d and f block are in between s and p block elements. Their properties show a transition in systematically and they vary from left to right.

Q.5 What is law of octaves?

Ans. Newland in 1864 classified 62 elements on the basis of their increasing atomic masses. According to Newland, "every eighth element has the properties similar to the first one if they are arranged according to increasing atomic masses".

Q.6 What is Mendleev's periodic law?

Ans. If the elements are arranged according to their ascending atomic masses, they show periodicity (repitition of physical and chemical properties after regular intervals) in the properties. It is called periodic law.

Q.7 What is modern periodic law?

Ans. According to Mosley (1913), if elements are arranged according to their ascending order of atomic numbers then chemical properties repeat in the periodic manner. It is called modern periodic law.

Q.8 Why the size of cation is smaller than its parent element

Ans. (a) When electron is removed from the valence shell there is increase in the nuclear charge.

(b) The size of atom reduces. For example; covalent radius of Na 157 Pm ionic radian of Na^+ is 95 Pm (Atomic radius of Na = 186 Pm).

Q.9 Na^+, Mg^{+2}, Al^{+3} are iso electronic in nature. Do these ions have same ionic radii?

Ans. Na^+, Mg^{+2}, Al^{+3} all have 10 electrons in their shells but their ionic radii are not same. Ionic radii decrease due to increase in the nuclear charge from left to right. Na^+ has 11 P, Mg^{+2} has 12 P and Al^{+3} has 13 protons in it.

Q.10 Why second ionization energy is greater than first ionization energy?

Ans. The amount of energy required to remove an electron from an atom in the gaseous state to produce gaseous positive ion is called first ionization energy.

The amount of energy required to remove second electron from mono-positive ion is called second ionization energy. Second ionization energy is greater than first ionization energy due to increase in the nuclear charge. For example:

$$Mg \longrightarrow Mg^{+1} + 1\bar{e} \qquad (I.E = 738 \text{ kJ/mol})$$

$$Mg^{+1} \longrightarrow Mg^{+2} + 1\bar{e} \qquad (I.E = 1451 \text{ kJ/mol})$$

Q.11 What is the trend in the metallic character in the periodic table?

Ans. Metallic character increases down the group due to increase in the atomic radii. The metallic character decreases along the period. Elements on the left side of the periodic table are metal and on the right side are non-metals. Out of all the elements known so far 82 are metals.

Q.12 What is the trend of melting point of elements in the periodic table?

Ans. Melting point of elements decreases down the group due to increases in the atomic size. Melting point of halogens increases down the group due to increased polarizibility of the molecule. I_2 has highest melting point among the halogens. Melting point of elements increases upto the middle (IV group) in the period and then decreases.

Q.13 What is the maximum oxidation state of an element?

Ans. The maximum oxidation state of an element is equal to its group number in the periodic table. For example Cl has maximum oxidation state of +7 in $KClO_4$ because it belongs to group VII.

Nitrogen in the V group has maximum oxdiation state of +5.

O_2 and F_2 are exceptions.

Q.14 Which oxide has greater acidic character MnO or Mn_2O_7?

Ans. Acidic character increases by increasing the oxidation number of the metal. Oxidation number of Mn in MnO is +2 while in Mn_2O_7 it is +7.

Therefore, Mn_2O_7 is more acidic than MnO. MnO is more basic in nature.

Similarly FeO is more basic than Fe_2O_3.

Q.15 Why second electron affinity has positive sign?

Ans. The amount of energy evolved or absorbed when an electron is added to the outermost shell of an atom in the gaseous state is called electron affinity. When an electron is added to an atom it form negative ion. The second electron is added into the negative ion. There will be repulsion between the electrons. To overcome the repulsive force, energy will be required. Therefore, second electron affinity is always positive.

$$O + 1\bar{e} \longrightarrow O^- \qquad (EA = -141 \text{ kJ/mol})$$

$$O^- + 1\bar{e} \longrightarrow O^{-2} \qquad (EA = +780 \text{ kJ/mol})$$

Q.16 What is the trend in ionic character of halides in the periods of the periodic table?

Ans. Along the period, ionic character decreases due to the decrease in the difference of electronegativity. For example; in the third period, NaCl and $MgCl_2$ are ionic, $AlCl_3$ is polymeric, and PCl_3, S_2Cl_2, $SiCl_4$ are covalent in nature.

Q.17 What is hydration energy?

Ans. The amount of heat evolved or absorbed when 1 mol of gaseous ions dissolve in water to give infinitely dilute solution. Its units are kJ/mol. For example:

$$H^+_{(g)} + \text{Water} \longrightarrow H^+_{(aq)} \qquad (\neg H = -1075 \text{ KJ/mol}^{-1})$$

Factors affecting Hydration Energies:

(i) Size of Ions: Greater the size of the ions, smaller the charge density and smaller the hydration energy.

(ii) Amount of Charge: Greater the amount of the charge, greater the hydration energy.

(iii) Hydration energy of negative ions is smaller.

Variations in the Periods and Groups:

(i) Hydration energy increases along the period for iso electronic ions in case Na^{+}, Mg^{+2}, Al^{+3}. These are iso electronic. The hydration energy increases with the increasing positive charge.

(ii) Hydration energy decreases down the group. For example: Li^{+} has greater hydration energy than Cs^{+} because Li^{+} has positive charge concentrated on a smaller volume relative to Cs^{+} where the charge rests on a bigger volume and is more diffused.