CHEMISTRY OF NON-TRANSITION ELEMENT

The non-transition elements are the main group elements that are 's' and 'p' block elements.

S-BLOCK :-

Alkali metals :-

These belong to group I of s-block of the modern periodic table. Group I elements include, lithium ($_{3}Li$), sodium ($_{11}Na$), potassium ($_{19}K$), rubidium ($_{37}Rb$), cesium ($_{55}Cs$) and francium ($_{87}Fr$).

These elements are called alkali metals because their hydroxides are very strong alkalies.

Electronic Configuration :-

The general outer electronic configurat^{n} of alkali metals is 'ns^1'.

$_{3}Li$ – $[He]\ 2s^1$

$_{11}Na$ – $[Ne]\ 3s^1$

$_{19}K$ – $[Ar]\ 4s^1$

$_{37}Rb$ – $[Kr]\ 5s^1$

$_{55}Cs$ – $[Xe]\ 6s^1$

$_{87}Fr$ – $[Rn]\ 7s^1$

Atomic and ionic radii :-

Within the group, the atomic as well as ionic radii of alkali metals increase with increase in atomic number. It is because of the increase in the no. of energy shells. e.g., the size of Na atom (having 3 shells) is Larger than size of Lithium atom (having 2 shells).

(2) Ionisation potential (I.P) / Ionisation energy (I.E) :-

On moving down the group, the ionisation potential decreases. It is because the s-electrons of outermost shell are getting away from the nucleus due to the addition of new shells. Therefore, the attractive pull exerted by the nucleus on the outermost electron decreases. Hence, electron can be removed by expending less energy.

Oxidation state or valency :-

All alkali metals form unipositive ions (+1).

Reason :-

Alkali metal lose an electron from the outermost 's' orbital and attain stable noble gas configuration. Hence, these metal form univalent positive ions.

$$M(g) \longrightarrow M^{+}(g) + e^{-}(g)$$

(Alkali metal ion — stable inert gas configuration)

Electronegativity and electron affinity :-

On moving down the group, the electronegativity and electron affinity of the alkali metals decreases. It is due to the increase in size and hence their nuclear charge decreases.

(3) Alkaline earth metals:~

These belong to group 2 of s-block of the modern periodic table.

Group 2 elements include beryllium ($_{4}Be$), magnesium ($_{12}Mg$), calcium ($_{20}Ca$), strontium ($_{38}Sr$), barium ($_{56}Ba$) and radium ($_{86}Ra$).

These were named alkaline earths because:
(i) these were alkaline in nature like alkali metal oxides.

(ii) these were found in earth surface.

Electronic configuration:~

The general outer electronic configuration of alkali metals is 'ns^2'.

$_{4}Be$ - $[He]\ 2s^2$

$_{12}Mg$ - $[Ne]\ 3s^2$

$_{20}Ca$ - $[Ar]\ 4s^2$

$_{38}Sr$ - $[Kr]\ 5s^2$

$_{56}Ba$ - $[Xe]\ 6s^2$

$_{88}Ra$ - $[Rn]\ 7s^2$

Atomic and ionic radii:~

Within the group, the atomic as well as ionic radii of alkaline earth metals increase with the increase in the number of energy shells. For example, Mg atom (having 3 shells) is larger than the size of beryllium atom (having 2 shells).

(1) Ionisation potential :-

On moving down the group, the ionisation potential decreases. It is because the s-electrons of the outermost shell are getting away from the nucleus due to the addition of new shells. Therefore, the attractive pull exerted by the nucleus on outermost electron decreases. Hence, electron can be removed by expending less energy.

Alkaline earth metals have smaller size and greater value of nuclear charge as compared to the corresponding alkali metals. Therefore, the electrons in alkaline earth metals are more firmly held than in the alkali metals. Hence, ionisation energy of alkaline earth metals is greater than the corresponding alkali metals.

Oxidation state or Valency :-

All alkaline earth metals form bipositive ions (+2).

Reason:-

Alkaline earth metal lose two electrons from the outermost 's' orbital and attain noble gas configuration. Hence, these metal form bipositive ions.

$$M(g) \longrightarrow M^{2+} + 2e^-$$

(Alkaline earth metal ion — Stable inert gas configuration)

Electronegativity and electron affinity :-

On moving down the group, the electronegativity and electron affinity of the alkaline earth metals decreases. It is due to the increase in size and hence their nuclear charge decreases

⑤ **P-BLOCK :~**

The p-block elements are found on the right side of the periodic table. They include the boron, carbon, nitrogen, oxygen and fluorine families in the addition to the noble gases.

BORON FAMILY :~

The boron group are the chemical elements in group 13 of the periodic table, comprising boron (${}_{5}B$), aluminium (${}_{13}Al$), gallium (${}_{31}Ga$), indium (${}_{49}In$), thallium (${}_{81}Tl$)

Electronic configuration :~

The general outer electronic configuration of Boron family is '$ns^2 np^1$'.

${}_{5}B - [{}_{2}He]\ 2s^2\ 2p^1$

${}_{13}Al - [{}_{10}Ne]\ 3s^2\ 3p^1$

${}_{31}Ga - [{}_{18}Ar]\ 3d^{10}\ 4s^2\ 4p^1$

${}_{49}In - [{}_{36}Kr]\ 3d^{10}\ 5s^2\ 5p^1$

${}_{81}Tl - [{}_{54}Xe]\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^1$

Oxidation state or valency :~

* Boron and Aluminium forms +3 oxidation states.
* Gallium, indium and thallium show +3 as well as +1. The +1 oxidation show due inert pair effect.

The stability of +1 oxidation state increases. While that of +3 oxidation state decreases down the group due to increase in inert pair effect down the group.

i.e.,

(i) Decreasing order of stability of +1 oxidation state is $Tl^{+} > In^{+} > Ga^{+}$

⑥ (ii) Decreasing order of stability of +3 oxidation states is $Ga^{3+} > In^{3+} > Tl^{3+}$

Atomic and ionic radius :-
The atomic radii of these elements do not show regular trend down the group.

i.e Tl > In > Al > Ga > B

Reason :-
When the effective nuclear charge is greater than screening effect, the atomic radius decreases or vice versa.

Ionisation energy :-
The ionisation energy of these elements do not show regular trend down the group.

i.e, B > Tl > Ga > Al > In

Reason :-
When the effective nuclear charge is greater than screening effect, the ionisation energy increases or vice-versa

Electronegativity :-
The electronegativity of these elements do not show regular trend down the group.

i.e., B > Tl > In > Ga > Al

Reason :-
When the effective nuclear charge is greater than screening effect, the electronegativity increases or vice-versa.

⑦ Carbon family:-

The carbon group are the chemical elements in group 14 of the periodic table, comprising carbon ($_{6}C$), silicon ($_{14}Si$), Germanium ($_{32}Ge$), stannum/tin ($_{50}Sn$) and Lead ($_{82}Pb$).

Electronic configuration:- The general outer electronic configuration of carbon family is $ns^2 np^2$,

$_{6}C - [He]^{2}\ 2s^2\ 2p^2$

$_{14}Si - [Ne]^{10}\ 3s^2\ 3p^2$

$_{32}Ge - [Ar]^{18}\ 3d^{10}\ 4s^2\ 4p^2$

$_{50}Sn - [Kr]^{36}\ 4d^{10}\ 5s^2\ 5p^2$

$_{82}Pb - [Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^2$

Oxidation state / valency:-

All group 14 elements form +4 oxidation state because group-14 elements lose 4 electrons from the outer shell and attain noble gas configuration.

And also Ge, Sn and Pb shows +2 oxidation state due to inert pair effect

Thus, the decreasing order of stability of +2 and +4 oxidation state is:-

$Pb^{+2} > Sn^{2+} > Ge^{2+}$

$Ge^{4+} > Sn^{4+} > Pb^{+4}$

(8) Ionisation energy :-

The ionisation energy of these elements do not show regular trend down the group.

i.e., C > Si > Ge > Pb > Sn

Reason :-

When the effective nuclear charge greater than screening effect, the ionisation energy increases or vice-versa.

Atomic radius :-

The atomic radii of these elements don't show regular trend down the group.

i.e., Pb > Sn > Ge > Si > C

Reason :-

When the effective nuclear charge greater than screening effect, the atomic radius decreases or vice-versa.

Electronegativity :-

The electronegativity of these elements don't show regular trend down the group.

i.e.,

Reason :-

When the effective nuclear charge greater than screening effect, the electronegativity increases or vice-versa.

(9) **Nitrogen family:-**

The nitrogen family are the chemical elements in group 15 of the periodic table, comprising nitrogen (N), phosphorus (P), Arsenic (As), Antimony (Sb) and Bismuth (Bi).

Electronic configuration:-

The general outer electronic configuration of nitrogen family is '$ns^2 np^3$'.

$_{7}N - [He]^{2}\ 2s^2\ 2p^3$

$_{15}P - [Ne]^{10}\ 3s^2\ 3p^3$

$_{33}As - [Ar]^{18}\ 3d^{10}\ 4s^2\ 4p^3$

$_{51}Sb - [Kr]^{36}\ 4d^{10}\ 5s^2\ 5p^3$

$_{83}Bi - [Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^3$

Oxidation no. or valency:-

+3 and +5 oxidation states are shown by almost all elements. +3 oxidation show due to the inert pair effect and +5 oxidation show due to the loss of five electrons to attain noble gas configuration. As we move down the group the inert pair effect increases regularly.

N also show -3 oxidn state due to the gain 3e⁻s. e.g. NH_3

Atomic radii:-

The atomic and ionic radii of group-15 elements are smaller than the atomic radii of the corresponding group-14 elements. This is because of increase in nuclear charge. However, on moving down the group the atomic radii increases due to increase in no. of shells or orbits.

i.e., Bi > Sb > As > P > N

(10) Ionisation energy :-

The ionisation energy of group-15 elements are higher than the corresponding group-14 elements due to the decrease in size (atomic radii). However, on moving down the group, the ionisation energy decreases due to the increase in size.

i.e., N > P > As > Sb > Bi

Electronegativity :-

The electronegativity of group-15 elements are higher than the corresponding group-14 elements due to decrease in size and it decreases on moving down the group.

i.e., N > P > As > Sb > Bi

(11) **Oxygen family :~(chalcogen)**

The oxygen family are the chemical elements in the group 16 of the periodic table, consist of oxygen ($_{8}O$), sulphur ($_{16}S$), selenium ($_{34}Se$), Tellurium ($_{52}Te$) and Polonium ($_{84}Po$).

Electronic configuration:- The outer electronic configuration of chalcogen is '$ns^2 np^4$'

$_{8}O - [He]^{2}\ 2s^2\ 2p^4$

$_{16}S - [Ne]^{10}\ 3s^2\ 3p^4$

$_{34}Se - [Ar]^{18}\ 3d^{10}\ 4s^2\ 4p^4$

$_{52}Te - [Kr]^{36}\ 4d^{10}\ 5s^2\ 5p^4$

$_{84}Po - [Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^4$

Oxidation no. or valency:-

+4 and +6 oxidation state are shown by more elements. +4 oxidation state show due to inert pair effect and +6 oxidat^{n} state show due to loss of six electrons to attain noble gas configuration.

O and S shows -2 oxidat^{n} states due to gain of 2 electrons e.g., H_2O and H_2S respectively.

As we move down the group, the inert pair effect increases regularly.

(12) **Atomic radii:-**

The atomic and ionic radii of group-16 elements are smaller than atomic radii of the corresponding group-15 elements This is because of increase of nuclear change. However, on moving down the group the atomic radii increases due to increase in no. of shells or Orbits.

i.e. Po > Te > Se > S > O

Ionisation energy:-

The ionisation energy of group 16 elements are higher than the corresponding group 15 elements due to decrease in size (atomic radii). However, on moving down the group the ionisation energy decreases due to increase in size.

i.e., O < S < Se < Te < Po

Electronegativity:-

The electronegativity of group 16 elements are higher than the corresponding group 15 elements due to decrease in size and it decreases on moving down the group.

O < S < Se < Te < Po

(13) HALOGEN FAMILY :-

The halogen family are the chemical elements in the group-17 of the periodic table, this family consist of Fluorine ($_{9}F$), Chlorine ($_{17}Cl$), Bromine ($_{35}Br$), Iodine ($_{53}I$) and Astatine ($_{85}At$).

Electronic configuration :-

The outer electronic configuration of Halogen family is '$ns^2\ np^5$'

$_{9}F - [_{2}He]\ 2s^2\ 2p^5$

$_{17}Cl - [_{10}Ne]\ 3s^2\ 3p^5$

$_{35}Br - [_{18}Ar]\ 3d^{10}\ 4s^2\ 4p^5$

$_{53}I - [_{36}Kr]\ 4d^{10}\ 5s^2\ 5p^5$

$_{85}At - [_{54}Xe]\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^5$

Oxidation no. or valency :-

-1, +5 and +7 Oxidation state are shown by more elements. -1 Oxidation state shows ↑ es due to the gain of one electron e.g., HF, HCl, HBr, HI etc. +5 oxidat? state show due the inert pair effect. As we move down the group, the inert pair effect increases regularly. +7 Oxidation state show due to the loss of 7 e⁻s to attain stable configuration.

(14) Atomic radii :-

The atomic and ionic radii of group 17 elements are smaller than atomic radii of the corresponding group-16 elements. This is because of increase of nuclear charge. However, on moving down the group the atomic radii increases due to increase in no. of shells or orbits.

i.e., At > I > Br > Cl > F

Ionisation energy :-

The ionisation energy of group 17 elements are higher than the corresponding group 16 elements due to decrease in size (atomic radii). However, on moving down the group, the ionisation energy decreases due to increase in size.

i.e., F > Cl > Br > I > At

Electronegativity :-

The electronegativity of group 17 elements are higher than the corresponding group 16 elements due to decrease in size and it decreases on moving down the group.

i.e., F > Cl > Br > I > At

(15) **Noble Gases :- (Group-18 elements)**

The elements of zero group are called noble gases or inert gases. The element belonging to this group are helium ($_{2}He$), neon ($_{10}Ne$), argon ($_{18}Ar$), Krypton ($_{36}Kr$), Xexon ($_{54}Xe$) and Radon ($_{86}Rn$).

These are called inert gases because they are chemical inert (inactive). It means they don't take part actively in any chemical reactions at all at ordinary temp. It is further because, they have a stable no. of electrons (doublet in He and octate for others) in their outermost shell. Hence, these are also called monoatomic gases.

Electronic configuration :-

Except helium which has $1s^2$ electronic configuration, all other elements of the noble gas family have '$ns^2\ np^6$' configuration.

$_{2}He - 1s^2$

$_{10}Ne - [He]\ 2s^2\ 2p^6$

$_{18}Ar - [Ne]\ 3s^2\ 3p^6$

$_{36}Kr - [Ar]\ 3d^{10}\ 4s^2\ 4p^6$

$_{54}Xe - [Kr]\ 4d^{10}\ 5s^2\ 5p^6$

$_{86}Rn - [Xe]\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^6$

Atomic radii :-

On moving down the group i.e., from 'He' to 'Rn' the atomic radii increases due to the addition of new energy shell.

(16) **Ionisation energy:-**

Due to the stable electronic configuration noble gases have maximum ionisation energy along a particular period. On moving down the group i.e, from 'He' to 'Rn', the ionisation energy decrease due to increase in atomic size.

Electron affinity:-

Since the noble gases have stable electronic configuration, therefore they have no tendency to accept electrons. Hence all noble gases have zero electron affinity.