

Periodic Classification of Elements :-

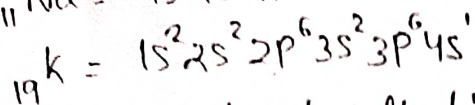
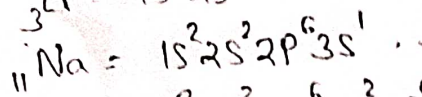
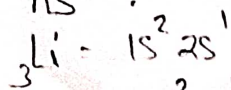
Electronic basis of periodic classification of elements

The method of arranging similar element in one group or separating them from dissimilar elements is called classification of elements.

Arrangements of elements with similar properties after certain intervals when these are arranged in the increasing order of their atomic number is called periodicity. The cause of periodicity ^{of} similar ~~is~~ elements repeat in regular interval is called electronic configuration.

Ex Alkali metal [Li, Na, K, Rb, Cs, Fr]

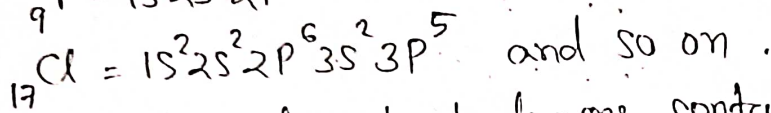
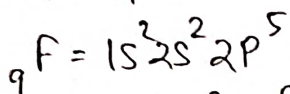
The general outer electronic configuration is ns^1 .



All elements of alkali metals contains one electron in their valance shell.

Ex 2 Consider halogen elements: F, Cl, Br, ~~I~~, At.

The general outer electronic configuration is $ns^2 np^5$.



All elements of halogens contains 7 numbers of valence electron.

From the above two examples it is clear that the properties of element ^{get} repeated after intervals 2, 8, 8, 18, 18, 32, because similar electronic configuration, elements repeat after these interval. The no. 2, 8, 8, 18, 18, 32 are called magic number.

Ex $H = 1$

$$\text{Li} = 1 + 2 = 3$$

$$\text{Na} = 3 + 8 = 11$$

$$\text{K} = 11 + 8 = 19$$

$$\text{Rb} = 19 + 18 = 37$$

$$\text{Cs} = 37 + 18 = 55$$

$$\text{Fr} = 55 + 32 = 87$$

Ex $0 = 8$

$$\text{S} = 8 + 8 = 16$$

$$\text{Se} = 16 + 18 = 34$$

$$\text{Te} = 34 + 18 = 52$$

$$\text{Po} = 52 + 32 = 84$$

Classification of elements into blocks -

On the basis of electronic configuration the incompletely filled orbitals, the elements have been divided into three types.

(1) Representative or s-p block element

(2) Transition or d-block element

③ Inner transition or f-block element
Representative element or s-block element

elements in which their outermost shell are incomplete.
These are two types.

① S-block elements

General outer electronic configuration = $ns^{1 \text{ to } 2}$

Groups are IA & IIA

IA - ns^1 configuration = Alkali metals except hydrogen (H)

IIA - ns^2 configuration = Alkaline earth metal

Characteristics.

- (i) Soft metals having low b.p. & m.p.
- (ii) They have low Ionisation Energy (I.E).
- (iii) They are highly reactive to form univalent or divalent cation.
- (iv) Their salt impart colour to the flame.

Ex Na - Golden yellow colour

Sr - Crimson red colour

Ca - Brick red colour

Ba - Pea green colour

K - Violet colour

(v) In solid state, these are good conductor of heat & electricity.

(vi) Cations are diamagnetic (Not attracted by the magnetic field) and are colourless due to absence of unpaired electrons.

(vii) These are good reducing agent due to low I.E.

(viii) (IA are more than IIA)

(ix) Their hydroxides are strong bases.

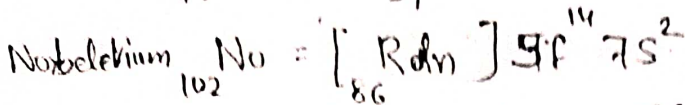
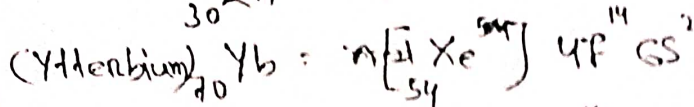
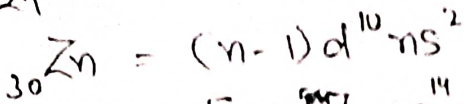
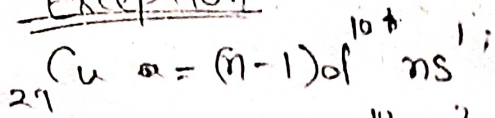
Ex NaOH, KOH, Ca(OH)_2 (lime water) etc.

② P-block elements

Their outer electronic configuration is $ns^2 np^{1 \text{ to } 6}$. present extreme right of the periodic table. Include groups = IIIA to VIIA - p zero group.

Their $(n-1)d$ orbital along with $(n-1)s$, $(n-1)p$ & ns orbitals are fully filled. The differentiating electron enters in np orbital.

Exception



included in the representative element

Cu & Zn = transition element more than representative element

Yb (Lanthanide) & No (actinides) more than representative element

Characteristics

- (i) Include both metal and non-metals. The metallic to non-metallic character increases from left to right across a period.
- (ii) They have high I.E increases from left to right.
- (iii) These are mostly ~~form~~ ^{form} covalent compounds.
- (iv) Reducing to oxidising character increases from left to right.
- (v) These are generally bad conductor of heat & electricity (except metals).
- (vi) Atomic radii decreases across a period & increases from top to bottom in a group.
- (vii) Heavier element shows variable valency.
- (viii) Elements of second period are known as bridge element. [diagonal relationship]
e.g. Li has similarities with Na & Mg.

Transition or d-block element

The general outer electronic configuration is $ns^2 (n-1)d^{1-10}$

There are 4 transition series.
1st transition series

(1) First transition series
Include 10 elements from atomic no. 21 to 30^{Zn}. The last electron enters in 3d orbital.

(2) Second transition series
Include 10 elements from atomic no. Yt (Yttrium) to 48^{Cd}. The last electron enters in 4d orbital.

(3) Third transition series
Include 10 elements from atomic number 57^{La} (Lanthanum) to 72^{Hf} (Hafnium). Then 80^{Hg} (Mercury).

(4) Fourth transition series
Contains 10 elements from 89^{Ac} (Actinium) to 104^{Rf} (Rutherfordium) to 112^{Uub} (Ununbium).

All d-block elements include the following groups: III B to VIII B or 3 to 7, VIII (8, 9, 10), IB & IIB (11 & 12). These are called transition elements because their properties are intermediate between s & p block elements. They are called d-block elements because the penultimate shell of electrons is expanded from 8 to 18 by the addition of d-orbitals.

Characteristics of d-block elements:-

(i) Elements are metals having high m.p. & b.p. due to relatively lower energy & presence of valence electrons.

(ii) Good conductor of heat & electricity.

(iii) Elements are paramagnetic (attracted by the magnetic field).

(iv) Generally form coloured compounds.

(v) Strong tendency to form complexes.

(vi) Elements like Mn, Co, Ni, Cu, V, and their compounds used as catalysts.

(vii) Elements exhibit variable oxidation states.

(viii) Inner transition elements or f-block elements

These are called inner transition elements because they have not only an incomplete d-subshell but also an incomplete f-subshell. These are called rare earth elements because they are less available. They are called f-block elements because

their antepenultimate orbit ($n-2f$ subshell being expanded from 18 to 32. There are two series.

(1) Lanthanide Series

At (58 Ce to 71 Lu)

Here $4f$ orbital is progressively filled up.
General outer electronic configuration $4f^{1-14} 5d^1 6s^2$

(2) Actinide series

(90 Th to 103 Lr)

Here $5f$ orbital progressively filled up.
The general outer electronic configuration is

$5f^{1-14} 6d^1 7s^2$

Characteristics of f -block Element :-

- (i) Heavy metals having high M.P & B.P.
- (ii) Show variable oxidation state.
- (iii) They form colour ion.
- (iv) They are paramagnetic due to presence of unpaired electron.
- (v) Actinides are radioactive in nature.

Atomic Radius

Atomic radius is defined as the distance from the centre of the nucleus up to its outermost electron is extended. Its unit is Angstrom (\AA).

Atomic radius decreases across a period from left to right due to high value of nuclear charge. And this value increases from top to bottom in a group due to successive addition of new orbitals.

Ionic Radius

It is the distance from the nucleus upto which it has influenced in an ionic bond. Its unit is Angstrom.

$$1 \text{\AA} = 10^{-8} \text{ cm} = 10^{-10} \text{ m}$$

- (1) The size of the cation smaller than the its corresponding neutral atom due to high value of nuclear charge and in cations number of orbitals also less than neutral atom.

Radius of $\text{Na}^+ < \text{Na}$

(ii) The size of anion is always larger than its corresponding neutral atom due to more number of electron in neutral atom and magnitude of nuclear charge is same.

(iii) From left to right across a period size of isoelectronic ion \rightarrow greater the nuclear charge of ion, greater will be the force of attraction, for the same number of electron. As a result, size of the ion decreases.

Ex N^{3-} , O^{2-} , F^- , Na^+ , Mg^{2+} , Al^{3+} are isoelectronic (nitride)

Order of \uparrow ion size of isoelectronic ion: $\text{N}^{3-} > \text{O}^{2-} > \text{F}^- > \text{Na}^+ > \text{Mg}^{2+} > \text{Al}^{3+}$

Covalent radius

Homonuclear molecules like H_2 , Cl_2 etc., The covalent radius of an atom is equal to the half of the distance between the centre of nuclei of two bonded atoms in a covalent bond.

N.B In a period, from left to right atomic radius goes on decreasing but ionic radius goes on decreasing.

Ionisation Potential

It is defined as the amount of energy required to remove one valance electron from an isolated neutral gaseous atom to form a monovalent gaseous +ve ion. Unit is eV (electron volt) or kilocalori/mole or kJ/mole

$$1 \text{ eV} = 23.06 \text{ kilocal/mole} = 96.5 \text{ kJ/mole}$$

Ex for Na, 1st ionisation energy (I.E) = 496 kJ/mole

Factor which influence the Ionisation Energy

- (i) I.E. increases with increase in nuclear charge.
- (ii) I.E. decreases with increase in atomic size.
- (iii) Greater is the screening effect, greater is the I.E. The screening effect order is in different orbital is $s > p > d > f$

(iv) Element in which an electron removed from s-orbital will have higher I.E. from p-orbital.

(v) I.E. of stable configuration (Noble gas configuration, half filled & full filled, outer orbitals) are greater than other configuration.

Trend in I.E. in a period: (second)

Li, Be > B, C, N > O, F, Ne

(i) I.E. of Be is greater than Li. Due to greater nuclear charge of Be, ^{2s} subshell is full filled and is paired, (1s²2s²)

* (ii) I.E. of Be is greater than B. Due to stable configuration of Be (1s²2s²).

(a) s-orbital penetrate more than p-orbital.

(b) Full filled 2s-orbital.

(iii) I.E. of C is more than B. Due to high value of nuclear charge of carbon.

(iv) I.E. of N is more than C.

Due to high value of nuclear charge & stable configuration of nitrogen atom. (Here outer p-orbital is exactly half filled).

(v) I.E. of N is greater than Oxygen.

Due to stable configuration of Nitrogen atom and paired electron present in Oxygen atom.

(vi) I.E. of F is more than Oxygen.

Due to more nuclear charge of F atom.

(vii) I.E. of Ne is greater than F.

Due to greater nuclear charge of Ne atom, stable configuration, pairing electron present in Ne atom.

N.B Ionisation energy of element in a group from top to bottom decreases due to increase in atomic size and screening effect.

ELECTRON AFFINITY

It is the amount of energy released when an extra electron is added to a neutral gaseous atom of an element to form a uni-negative gaseous ion is called Electron affinity.

Unit: eV, kcal/mole; kJ/mole

Electron affinity depends upon the following points

- (i) Smaller the size of the atom, greater will be the electron affinity.
- (ii) Greater is the magnitude of effective nuclear charge of an atom, greater will be the electron affinity.
- (iii) Stable configuration atom has less electron affinity.

Variation in a Group:

In a group, electron affinity of elements decreases due to increase in atomic size & nuclear charge. Small size atom of the first element of each group, the electron affinity value is less than the second element due to repulsion between the electron. Hence, the electron affinity of Fluorine is less than Chlorine.

Variation along a period:

Electron affinity of element increases from left to right across a period due to increasing effective nuclear charge and decrease in size of atom.

EX Halogens are high value of electron affinity (E.A). Electron affinity of Be & Mg, N, and noble gases are nearly 0 or +ve due to stable configuration.

ELECTRO NEGATIVITY

The tendency or power of an atom in a molecule to attract the shared paired electron towards itself is known as electro negativity (E.N).

This value increases across a period from left to right ^{due to increase in nuclear charge}. Fluorine is the most electronegative element. Electronegativity value from top to bottom in a group decreases.

Factors influenced the Electronegativity: [2/1/11]

- (i) Smaller size ^{atom} ~~electron~~ has high electronegativity.
- (ii) Anion are less electronegative than atom.
- (iii) Molecules having triple bond, the attachment atom is more electronegative.
- (iv) The effective nuclear charge decreases with increase in atomic number and the size of atom also increases. Hence, electronegativity value decreases from top to bottom in a group.