

3: Classification of Elements and Periodicity in Properties

1. Dobereiner's Triads (1829): Dobereiner classified the elements into groups of three elements with similar properties in such a manner so that the atomic weight of the middle element was the arithmetic mean of the other two like

Element Li Na K, Atomic weight 7 23 39

Mean of atomic masses = $(7 + 39) / 2 = 23$

2. Newland's Octaves (1864) (Law of Octaves): Newland states that when elements are arranged in order of increasing atomic masses, every eighth element has properties similar to the first just like in the musical note like

sa re ga ma pa dha ni

Li Be B C N O F

Na Mg Al Si P S Cl

3. Mendeleev's Periodic Table: "The physical and chemical properties of the elements are a periodic function of their atomic masses."

At the time of Mendeleev, only 63 elements were known. This Periodic Table is divided into seven horizontal rows (periods) and eight vertical columns (groups). Zero group was added later on in the modified Mendeleev's Periodic Table.

4. Modern periodic law: Properties of elements are the periodic function of their atomic number.

5. Modern Periodic Table: This table was prepared by Bohr and is based upon the electronic configuration of elements. The table consists of 18 vertical columns called groups and 7 horizontal rows called periods. In a group, elements having similar outer electronic configurations. The first period contains 2 elements. The subsequent periods consist of 8, 8, 18, 18 and 32 elements, respectively. The seventh period is incomplete and like the sixth period would have a theoretical maximum (on the basis of quantum numbers) of 32 elements. In this form of the Periodic Table, 14 elements of both sixth and seventh periods (lanthanoids and actinoids, respectively) are placed in separate panels at the bottom.

6. Electronic configuration and types of Elements (s-, p-, d-, f- block):

s-Block Elements : The elements of Group 1 (alkali metals) and Group 2 (alkaline earth metals), having ns^1 and ns^2 outermost electronic configuration.

p-Block Elements: Belonging to Group 13 to 18 and these together with the s-Block Elements are called the Representative Elements or Main Group Elements. The outermost electronic configuration varies from ns^2np^1 to ns^2np^6 in each period.

This is the only block which contains metal, non-metal and metalloids. Examples of metalloids are B, Si, Ge, As, Sb, Te and At.

d-Block Elements (Transition metals): Belonging to Group 3 to 12 in the centre of the Periodic Table. These are characterized by the filling of inner d orbitals by electrons and are therefore referred to as d-Block Elements. General outer electronic configuration is $(n-1) d^{1-10} ns^{0-2}$.
d-block elements are generally colored, paramagnetic and exhibit variable valency.

f-Block Elements (Inner transition metals): The two rows of elements at the bottom of the Periodic Table, called the Lanthanoids, Ce (Z = 58) – Lu (Z = 71) and Actinoids, Th (Z = 90) – Lr (Z = 103), their outer electronic configuration is $(n-2)f^{1-14} (n-1)d^{0-1} ns^2$. The last electron added to each element is filled in f- orbital.

7. Notation for IUPAC nomenclature of elements with Atomic Number above 100:

Digit	0	1	2	3	4	5	6	7	8	9
Root	nil	un	bi	tri	quad	pent	hex	sept	oct	enn
Abbreviation	n	u	b	t	q	p	h	s	o	e

The IUPAC names and symbols of elements with $Z > 100$ are

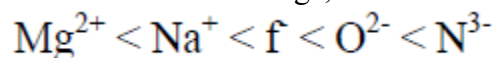
Z	101	102	103	104	105	106	107	108	109	110
IUPAC name	Unnilunium	Unnilbium	Unniltrium	Unnilquadium	Unnilpentium	Unnilhexium	Unnilseptium	Unniloctium	Unnilennium	Ununnilium
Symbol	Uun	Uub	Uut	Uuq	Uup	Uuh	Uus	Uuo	Uue	Uun

8. Atomic Radius: It is the distance from the centre of the nucleus to the outermost shell of electrons.

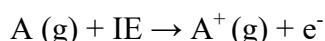
The atomic size decreases on moving from left to right in a period due to increase in effective nuclear charge and increases on moving from top to bottom in a group due to addition of new shells.

9. Ionic Radius: A cation is smaller than its parent atom because it has fewer electrons while its nuclear charge remains the same. The size of an anion will be larger than that of the parent atom because the addition of one or more electrons would result in increased repulsion among the electrons and a decrease in effective nuclear charge.

Isoelectronic ions: Higher the effective nuclear charge, smaller is the size like



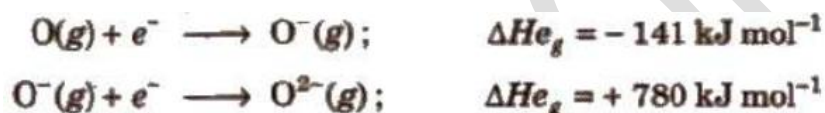
10. Ionization Enthalpy (IE): It is the amount of energy required to remove the loosely bound electron from the isolated gaseous atom. Its unit is kJ / mole



- > Factors which affecting IE are :(i) Atomic size: varies inversely (ii) Screening effect: varies inversely (iii) Nuclear charge: varies directly
- > Left to right in periods IE increases while down the group, it decreases.
- > IE values of inert gases are exceptionally higher due to stable configuration.
- > Successive ionization enthalpies - $IE_3 > IE_2 > IE_1$
- > Exception- (i) IE of N is greater than that of O due to stable half – filled 2p-orbitals.
- (ii) Among transition elements of 3d-series, 24Cr and 29Cu have higher IE_2 due to half-filled and fully-filled stable d-orbitals.

11. Electron Gain Enthalpy (EGE or ΔH_{eg}): It is the amount of energy released when an electron is added in an isolated gaseous atom.

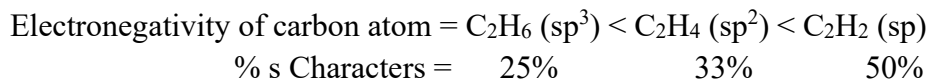
First electron gain enthalpy is negative while the other successive electron gain enthalpy will be positive due to repulsion between the electrons already present in the anion and the electron being added.



- > Factors which affecting electron gains enthalpy are:
 - (i) Atomic size: varies indirectly (ii) Nuclear charge: varies directly
- > Along a period, moving from left to right electron gain enthalpy increases in negative value while on moving down the group, it becomes decreases in negative value.
- > Noble gases have positive electron gain enthalpies.
- > Elements having half-filled and fully-filled orbitals exhibit more stability. Therefore, electron gain enthalpy will be low for such elements.
- > Electron gain enthalpy can be measured by Born-Haber cycle and elements with high ΔH_{eg} are good oxidizing agent.
- > Exception- F and O atom have lower electron gain enthalpy, than Cl and S respectively $Cl > F$; $S > O$ because when an electron is added to O or F, the added electron goes to the smaller $n = 2$ quantum level and suffers significant repulsion from the other electrons present in this level. For the $n = 3$ quantum level (S or Cl), the added electron occupies a larger region of space and the electron - electron repulsion is much less.

12. Electronegativity (EN): The tendency of an atom to attract the shared electron pair towards itself in a covalent bond. Various factors with which electronegativity varies are:

- > Factors which affecting EN are:
 - (i) Atomic size: varies inversely
 - (ii) Charge on the ion: varies directly, e.g., $Li < Li^+$, $Fe^{2+} < Fe^{3+}$
 - (iii) Hybridisation : % s-character in the hybrid orbital is more than EN should be more.

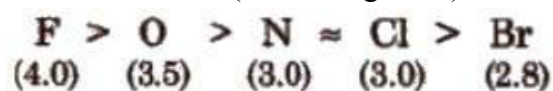


- > In periods as we move from left to right electro negativity increases, while in the groups

Electro negativity decreases down the group. For noble gases, its value is taken as zero.

> Electro negativity helps to predict the polarity of bonds and dipole moment of molecules.

> Electro negativity order of some elements (on Pauling scale) is



> Electronegativity of an element is a variable property. It is different in different compounds. For example the electronegativity of N is different in NH₃ and NO₂.

13. Valency: It is defined as the combining capacity of the element. The valency of an element is related to the electronic configuration of its atom and usually determined by unpaired electrons present in the valence shell,

> On moving along a period from left to right, valency increases from 1 to 4 and then decreases to zero (for noble gases) while on moving down a group the valency remains the same.

> Transition metals exhibit variable valency because they can use electron (from outer as well as penultimate shell).

Covalency:

14. Electropositive or Metallic character (Reducing Property): The tendency of an element to lose electrons and forms positive ions (cations) is called electropositive or metallic character.

> The elements having lower ionization energies have higher tendency to lose electrons, thus they are electropositive or metallic in their behavior.

> Alkali metals are the most highly electropositive elements.

Periodicity: It decreases from moving left to right in a period and increases from top to bottom in a group.

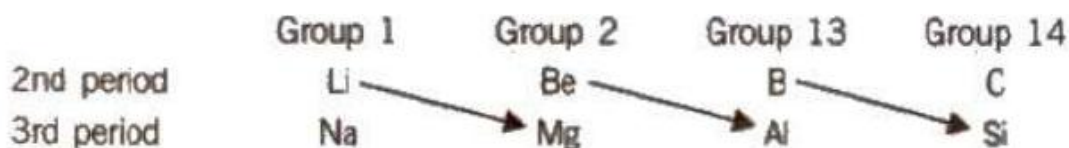
15. Electronegative or Non-Metallic character (Oxidizing Property): The tendency of an element to accept electrons to form an anion is called its non metallic or electronegative character. The elements having high electro-negativity have higher tendency to gain electrons and forms anion. So, the elements in the upper right hand portion of the periodic table are electro-negative or non-metallic in nature.

Periodicity: It increases from left to right in a period and decreases from top to bottom in a group.

Atomic Size↑; IE↓; EGE (-)↓; EN↓; Non Metallic Character↓; Metallic Character↑

16. Anomalous Properties of Second Period Elements (Diagonal Relationship):

Certain elements of 2nd period show similarity in properties with their diagonal elements in the 3rd period as shown below:



Thus, Li resembles Mg, Be resembles Al and B resembles Si. This is called **diagonal relationship**. Elements of third period, i.e., Mg, Al and Si are also known as bridge elements.

The anomalous behavior is due to their

- (a) Small size, large charge/ radius ratio and high electro negativity of the elements.
- (b) In addition, the first member of group has only four valence orbitals ($2s$ and $2p$) available for bonding, whereas the second member of the groups have nine valence orbitals ($3s$, $3p$, $3d$). As a consequence of this, the maximum covalency of the first member of each group is 4 (e.g., boron can only form $[\text{BF}_4]^-$, whereas the other members of the groups can expand their valence shell to accommodate more than four pairs of electrons e.g., aluminum forms $[\text{AlF}_6]^{3-}$).
- (c) Furthermore, the first member of p -block elements displays greater ability to form $p\pi - p\pi$ multiple bonds to itself (e.g., $\text{C} = \text{C}$, $\text{C} \equiv \text{C}$, $\text{N} = \text{N}$, $\text{N} \equiv \text{N}$) and to other second period elements (e.g., $\text{C} = \text{O}$, $\text{C} = \text{N}$, $\text{C} \equiv \text{N}$, $\text{N} = \text{O}$) compared to subsequent members of the same group.

17. Modern Periodic Table:

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
	s - Block												p - Block					
2	H																	He
8	Li (3)	Be											B	C	N	O	F	Ne
8	Na (11)	Mg	d - Block										Al	Si	P	S	Cl	Ar
18	K (19)	Ca	Sc (21)	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn (30)	3d	Ge	As		Br	Kr
18	Rb (37)	Sr	Y (39)										4d		Sb	Te	I	Xe
32	Cs (55)	Ba	La (57)										5d				At	Rn
32	Fr (87)	Ra	Ac (89)										6d					
f - Block																		
	Lanthanoids (4f)		Ce (58)															Lu (71)
	Actinoids (5f)		Th (90)															Lr (103)