Trigonometry

- Johann Dobereiner noticed that the middle element of each of the Triads had an atomic weight about half way between the atomic weights of the other two. Also the properties of the middle element were in between those of the other two members. Since Dobereiner's relationship, referred to as the Law of Triads, seemed to work only for a few elements, it was dismissed as coincidence.
- John Alexander Newlands in 1865 profounded the Law of Octaves. He arranged the elements in increasing order of their atomic weights and noted that every eighth element had properties similar to the first element.
- Periodic Law-The properties of the elements are a periodic function of their atomic weights.
- Modern Periodic Law-The physical and chemical properties of the elements are periodic functions of their atomic numbers.
- A modern version, the so-called "long form" of the Periodic Table of the elements, is the most convenient and widely used. The horizontal rows are called periods and the vertical columns, groups. Elements having similar outer electronic configurations in their atoms are arranged in vertical columns, referred to as groups or families.

- There are altogether seven periods. The period number corresponds to the highest principal quantum number (n) of the elements in the period. The first period contains 2 elements. The subsequent periods consists of 8, 8, 18, 18 and 32 elements, respectively.

- IUPAC-International Union of Pure and Applied Chemistry.
- IUPAC nomenclature for elements-

0 = nil	3 = tri	6 = hex	9 = enn
1 = un	4=quad	7=sept	
2 = bi	5=pent	8 = oct	

- The distribution of electrons into orbitals(s, p, d, f) of an atom is called its electronicconfiguration.

Atomic number	Symbol	Electronic configuration
3	Li	$1s^2 2s^1$ **(or)** $[He]2s^1$
11	Na	$1s^2 2s^2 2p^6 3s^1$ **(or)** $[Ne]3s^1$
19	K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ **(or)** $[Ar]4s^1$
37	Rb	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^1$ **(or)** $[Kr]5s^1$
55	Cs	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 6s^1$ **(or)** $[Xe]6s^1$
87	Fr	$[Rn]7s^1$

s-block elements

The elements of Group 1 (alkali metals) and Group 2 (alkaline earth metals) which have ns^1 and ns^2 outermost electronic configuration belong to the *s*-Block Elements. They are all reactive metals with low ionization enthalpies. They lose the outermost electron(s) readily to form 1+ ion (in the case of alkali metals) or 2+ ion (in the case of alkaline earth metals). The metallic character and the reactivity increase as we go down the group.

p-block elements

The *p*-Block Elements comprise those belonging to Group 13 to 18 and these together with the *s*-Block Elements are called the Representative Elements or Main GroupElements. At the end of each period is a noble gas element with a closed valence shell ns^2np^6 configuration. It is very difficult to alter this stable arrangement by the addition or removal of electrons, hence exhibit very low chemical reactivity. Preceding the noble gas family are two chemically important groups of non-metals. They are the halogens (Group 17) and the chalcogens(Group 16) have high negative electron gain enthalpies and readily add one or two electrons respectively to attain the stable noble gas configuration. The non-metallic character increases as we move from left to right across a period and metallic character increases as we go down the group.

d-block elements

These are the elements of Group 3 to 12 in the centre of the Periodic Table. These elements have the general outer electronic configuration (n-1) $d^{1-10}ns^{0-2}$ .transition metals form a bridge between the chemically active metals of *s*-block elements and the less active elements of Groups 13 and 14 and thus take their familiar name “TransitionElements”.

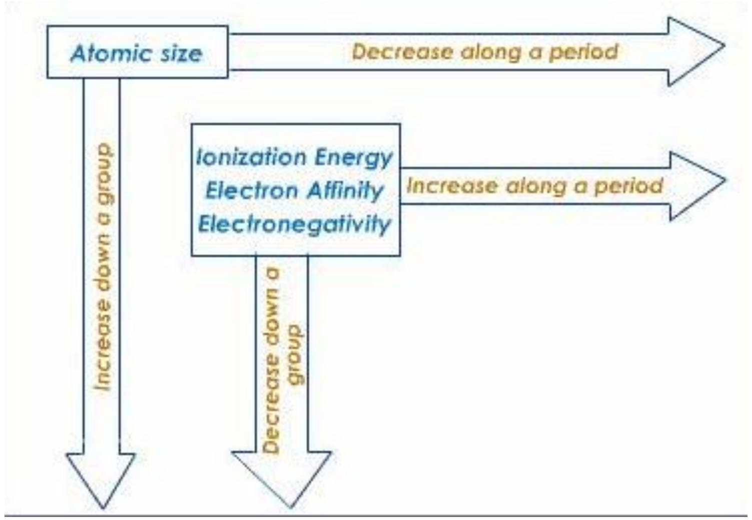
f-block elements

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) are characterised by the outer electronic configuration (n-2) f^{1-14} (n-1) $d^{0-1}ns^2$. The last electron added to each element is filled in *f*- orbital. These two series of elements are hence called the Inner-Transition Elements (*f*-Block Elements).

- The elements can be divided into Metals and Non-Metals. Metals comprise more than 78% of all known elements and appear on the left side of the PeriodicTable. Non-metals are located at the top right hand side of the Periodic Table.
- The atomic size generally decreases across a period. Within a family or vertical column of the periodic table, the atomic radius increases regularly with atomic number.
- The removal of an electron from an atom results in the formation of a cation, whereas gain of an electron leads to an anion. The ionic radii can be estimated by measuring the distances between cations and anions in ionic crystals.
- In general, the ionic radii of elements exhibit the same trend as the atomic radii.
- 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
- When we find some atoms and ions which contain the same number of electrons, we call them isoelectronic species. For example O2–, F–, Na+ and Mg2+ have the same number of electrons (10)
- A quantitative measure of the tendency of an element to lose electron is given by its Ionization Enthalpy. It represents the energy required to remove an electron from an isolated gaseous atom (X) in its ground state.
- The second ionization enthalpy will be higher than the first ionization enthalpy because it is more difficult to remove an electron from a positively charged ion than from a neutral atom.
- When an electron is added to a neutral gaseous atom (X) to convert it into a negative ion, the enthalpy change accompanying the process is defined as the Electron Gain Enthalpy.
- Group 17 elements have very high negative electron gain enthalpies because they can attain stable noble gas electronic configurations by picking up an electron.
- Noble gases have large positive electron gain enthalpies because the electron has to enter the next higher principal quantum level leading to a very unstable electronic configuration.

- Electron gain enthalpy becomes more negative with increase in the atomic number across a period. The effective nuclear charge increases from left to right across a period and consequently it will be easier to add an electron to a smaller atom since the added electron on an average would be closer to the positively charged nucleus.
- Electron gain enthalpy becomes less negative as we go down a group because the size of the atom increases and the added electron would be farther from the nucleus.
- **Anomaly**-Electron gain enthalpy of O or F is less than that of the succeeding element. This is 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.
- A qualitative measure of the ability of an atom in a chemical compound to attract shared electrons to itself is called electro-negativity.
- Electro-negativity generally increases across a period from left to right and decrease down a group from in the periodic table.
- The increase in electro-negativities across a period is accompanied by an increase in non-metallic properties of elements. Similarly, the decrease in electro-negativity down a group is accompanied by a decrease in non-metallic properties of elements.



Periodicity of Valence or Oxidation States

- The valence is the most characteristic property of the elements and can be understood in terms of their electronic configurations. The valence of representative elements is usually (though not necessarily) equal to the number of electrons in the outermost orbitals and / or equal to eight minus the number of outermost electrons.
- The oxidation state of an element in a particular compound can be defined as the charge acquired by its atom on the basis of electronegative consideration from other atoms in the molecule.

Anomalous Properties of Second Period Elements

- The first element of each of the groups 1 (lithium) and 2 (beryllium) and groups 13-17(boron to fluorine) differs in many respects from the other members of their respective group.
  For example, lithium unlike other alkali metals, and beryllium unlike other alkaline earth metals, form compounds with pronounced covalent character; the other members of these groups predominantly form ionic compounds.
- In fact the behaviour of lithium and beryllium is more similar with the second element of the following group i.e., magnesium and aluminium, respectively. This sort of similarity is commonly referred to as diagonal relationship in the periodic properties.

Sample Examples

- Show by a chemical reaction with water that Na_2O is a basic oxide and Cl_2O_7 is an acidic oxide.

  Solution

  Na_2O with water forms a strong basewhereas Cl_2O_7 forms strong acid.

  $Na_2O + H_2O \rightarrow 2NaOH$

  $Cl_2O_7 + H_2O \rightarrow 2HClO_4$

  Their basic or acidic nature can bequalitatively tested with litmus paper.

- Using the Periodic Table, predict theformulas of compounds which might beformed by the following pairs of elements;(a) silicon and bromine (b) aluminium andsulphur.

  Solution

  (a) Silicon is group 14 element with avalence of 4; bromine belongs to thehalogen family with a valence of 1.Hence the formula of the compoundformed would be $SiBr_4$.

  (b) Aluminium belongs to group 13 witha valence of 3; sulphur belongs togroup 16 elements with a valence of2. Hence, the formula of the compoundformed would be Al_2S_3.

- Which of the following species will havethe largest and the smallest size?Mg, Mg^{2+}, Al, Al^{3+}.

  Solution

  Atomic radii decrease across a period. Cations are smaller than their parentatoms. Among isoelectronic species, theone with the larger positive nuclear chargewill have a smaller radius.Hence the largest species is Mg; thesmallest one is Al^{3+}

- How would you justify the presence of 18elements in the 5th period of the PeriodicTable?

  Solution

  When n = 5, l = 0, 1, 2, 3. The order in which the energy of the available orbitals4d, 5s and 5p increases is 5s < 4d < 5p.The total number of orbitals available are9. The maximum number of electrons thatcan be accommodated is 18; and therefore18 elements are there in the 5th period.