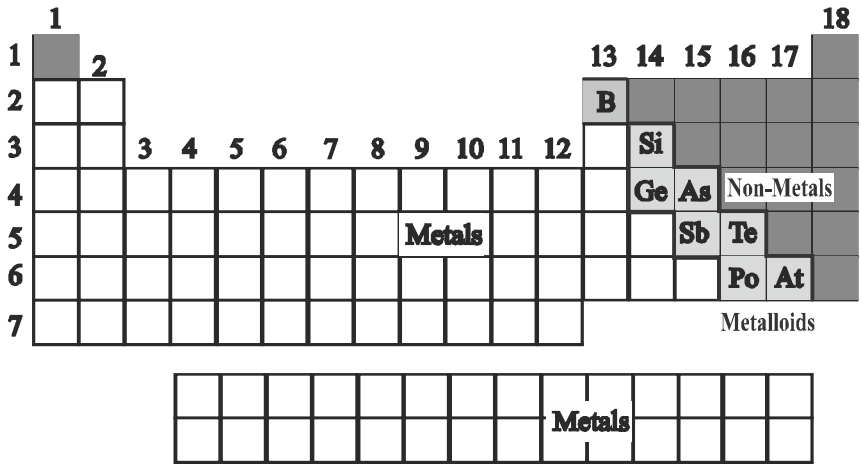
e.g. Si, As, Sb, Ge, Te



Position of metals, non-metals and metalloids in the periodic table.

Metals have the tendency to form cations by loss of electrons and this property make the elements as electro positive elements or metals.

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

- The tendency of an element to lose electron is closely connected to the (IE) of the element.
- Smaller the (IE) of an element, the greater will be its tendency to lose electrons and thus greater will be its metallic character.
- Greater the metallic character, greater the reducing nature.

* (IE) increases moving along a period left to right and decreases down the group, hence metallic and reducing nature decrease along the period and increase down the group.

Nuclear charge — increases left to right in period →

Shielding — increases left to right in period →

But the nuclear charge wins so :

Tendencey to form the ions ← decreases left to right in period

Reducing power ← decreases left to right in period

Metallic nature ← decreases left to right in period

Problem-2: Tendency to lose electrons shows reducing property of the element. Arrange the following in order of reducing property.

(a) Na,K,Rb (b) Na, Mg, Al (c) $F^{-}, Cl^{-}, Br^{-}, I^{-}$ (d) Mg, Ca, Sr

Soln. (a) $Na < K < Rb$ (b) $Na > Mg > Al$

(c) $F^{-} < Cl^{-} < Br^{-} < I^{-}$ (d) $Mg < Ca < Sr$

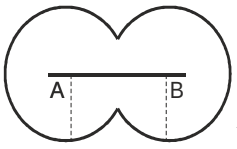
(iii) Size of atoms and ions

(a) Atomic Radii :

The distance from centre of the nucleus to the outermost shell is called radius of an atom, it depends on following factors:

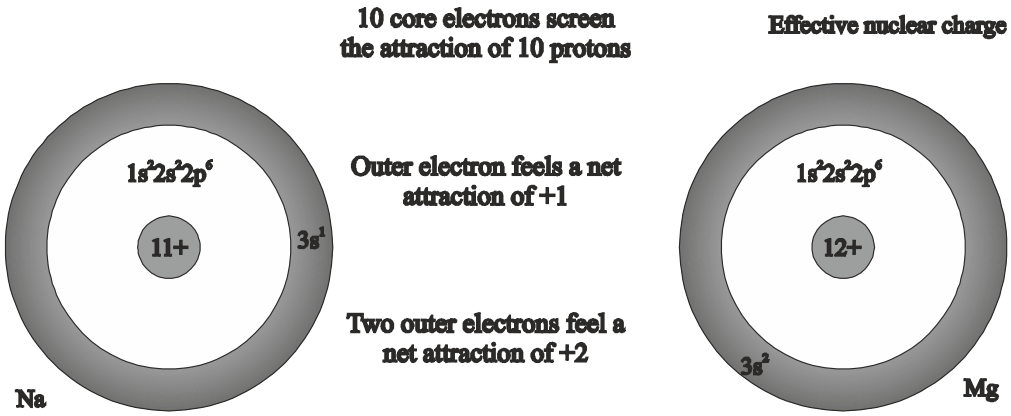
(i) Principal Quantum number (n)

(ii) Effective Nuclear Charge



$\frac{1}{2}AB$ = covalent radius

Effective nuclear charge (ENC) or Z*:

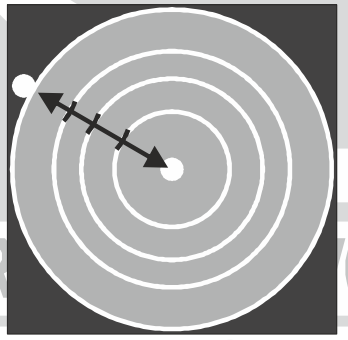


- The effective nuclear charge is the positive charge that an electron experience from the nucleus.
- It is the nuclear attractive force experienced by the electron when it is shielded by innerlying electrons.
- It is the nuclear charge reduced by shielding or screening from any intervening electrons.
- 2p electron is shielded more than or 2s electron as it penetrates the 1s orbital less than 2s orbital.
- As a result we have the energy sequence.

$2s < 2p$
$3s < 3p < 3d$
$4s < 4p < 4d < 4f$

As we move to atoms of elements of higher atomic number, the energy difference between orbitals of same value of n decreases.

Screening Effect:



The decrease in the nuclear force of attraction on the valence electrons or outermost electrons due to repulsive forces of inner lying electrons is called **screening effect.**

The electrons of inner shell repel the electrons of outermost shells. The electrons of outermost shell are thus shielded or screened from the nucleus by inner electrons.

As a result of screening effect the outermost electrons do not experience the complete nuclear charge. The radial distribution of electron densities indicates that 2s and 2p orbitals of H-atom have substantial values at the distance of 1s orbital also, the 2s and 2p orbitals penetrate the 1s orbtial.

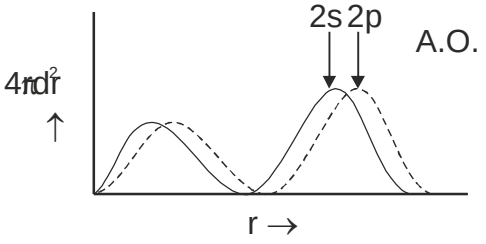


Figure: Probability distribution of electron density