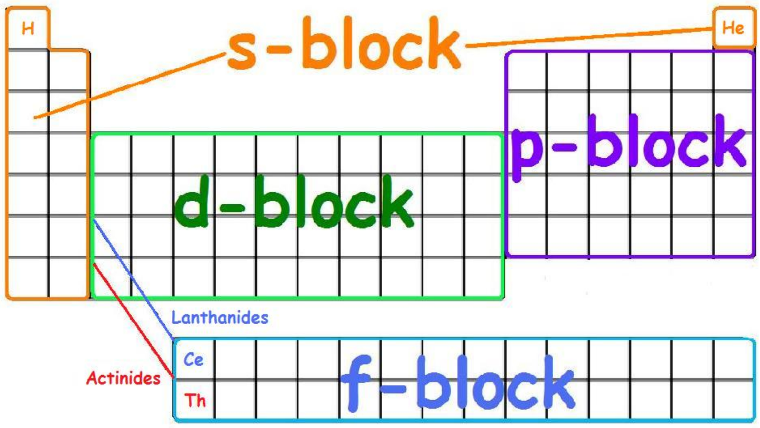
Modern Periodic Law:

- Properties of elements are the periodic function to their atomic numbers.
- The periodicity in properties is due to repetition of similar outer shell electronic configuration at a certain regular intervals.
- In modern periodic table is based on modern periodic law in which elements are arranged in increasing order of their atomic numbers.
- In the modern periodic table, the elements are arranged in rows and columns. These rows and columns are known as periods and groups respectively.
- The table consists of 7 periods and 18 groups
- Period indicates the value of 'n' (principal quantum number) for the outermost or valence shell.
- Same number of electrons is present in the outer orbitals (that is, similar valence shell electronic configuration

IUPAC Nomenclature for Elements with Atomic Number ? 100

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

Classification of Elements:



Elements:	Valence Shell Electronic Configuration	Nature	Position in Modern Periodic Table
s-block elements	ns^{1-2} (n = 1 to 7).	Metals	1 and 2 group elements
p – block elements	ns^2np^{1-6} (n = 2 to 7).	Metalloids & non metals but some of them are metals also.	groups 13 to 18
d-Block Elements	$(n-1)d^{1-10}\ ns^{1-2}$ (n = 4 to 7).	Metals	3 to 12 groups 3d series – Sc(21) to Zn (30) 4d series – Y (39) to Cd (48) 5d series – La (57), Hf (72) to Hg (80
f-Block Elements	$(n-2)f^{1-14}\ (n-1)s^2\ (n-1)p^6\ (n-1)d^{0-1}ns^2$ (n = 6 and 7).	Radioactive	group 3 4f series – Lanthanides – 14 Elements Ce (58) to Lu (71) 5f series – Actinides – 14 Elements Th (90) to Lw (103)

Periodicity in Atomic Properties:

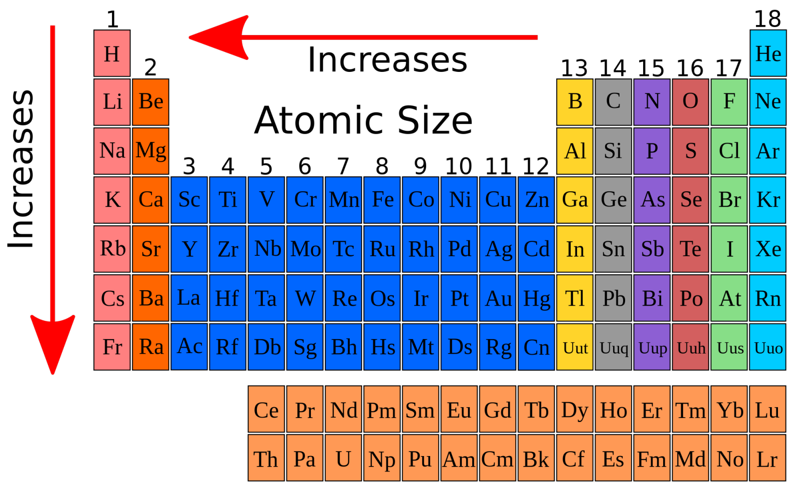
1. Atomic Radius:

- Within a given period atomic radius **decreases from left to right**. This is due to the effect of increase in nuclear charge while the electrons are being added to the same shell.
- Within a given group atomic radius increases down the group. This is due to the increase in

number of shells.

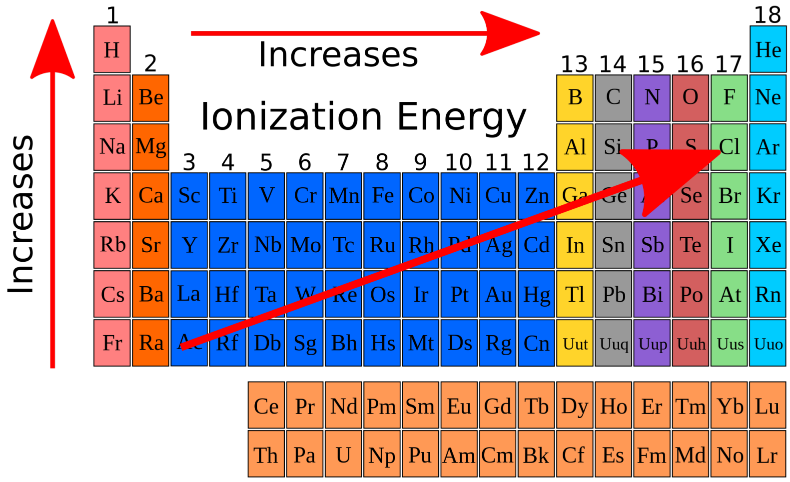
Period	Group							
	IA	IIA	IIIA	IVA	VA	VIA	VIIA	Zero
1.	H 0.37							He 0.93
2.	Li 1.34	Be 0.90	B 0.82	C 0.77	N 0.73	O 0.74	F 0.72	Ne 1.31
3.	Na 1.54	Mg 1.30	Al 1.18	Si 1.11	P 1.06	S 1.02	Cl 0.99	Ar 1.74

- In the first transition series the atomic size slightly decreases from Sc to Mn because effect of effective nuclear charge is stronger than the shielding effect. The atomic size from Fe to Ni remains almost the same because both the effects balance each other.
- The atomic size from Cu to Zn slightly increases because shielding effect is more than effective nuclear charge due to d^{10} structure of Cu and Zn.
- Inner transition elements – As we move along the lanthanide series, there is a decrease in atomic as well as ionic radius. The decrease in size is regular in ions but not so regular in atoms. This is called lanthanide contraction.



2) Ionisation potential or Ionisation Energy:

Ionization energy increases along the period while decreases down the group.



Factors which influence I.E.

- Atomic size: the larger the size of the atom, the smaller the I.E. i.e., I.E. μ

- Effective nuclear charge: The greater the effective charge on the nucleus of an atom, the more difficult it would be to remove an electron from the atom because electrostatic force of attraction between the nucleus and the outermost electron increases. So greater energy will be required to remove the electron.

- Penetration effect of orbitals: The order of energy required to remove electron from s,p,d-and ¦-orbitals of a shell is s>p>d>¦.

- Shielding or screening effect: Screening effect results in decrease of force of attraction between the nucleus and the outermost electron and lesser energy is required to separate the electron. Thus the value of I.P. decreases.

- Stability of half-filled and fully-filled orbitals: According to Hund's rule the stability of half filled or completely filled degenerate orbitals is comparatively high. So comparatively more energy is required to separate the electron from such atoms.

Successive Ionisation Energies

Element	IE_1	IE_2	IE_3	IE_4
Li	520.1	7297	11813	--
C	1086.2	2352	4620	6221
N	1402.1	2856	4577	7474

3) Electron Affinity:

Electron affinity increases along the period while decreases down the group.

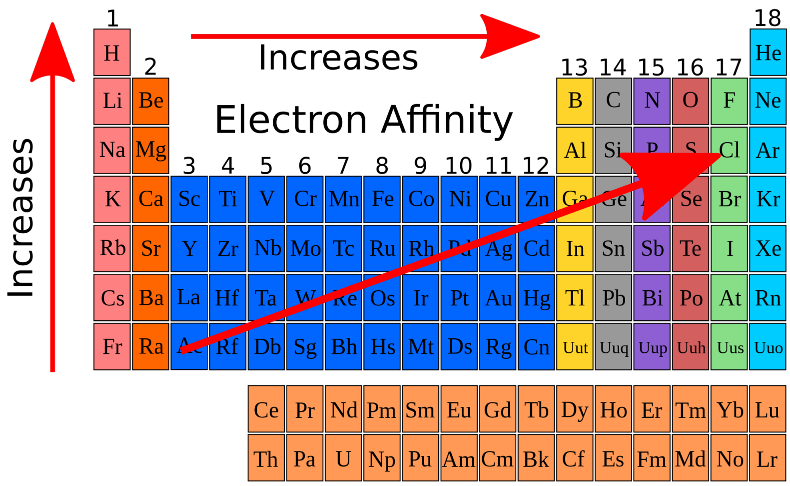
Factors affecting the magnitude of electron affinity

- Atomic size – In general electron affinity value decreases with the increasing atomic radius because electrostatic force of attraction decreases between the electron being added and the atomic nucleus due to increase of distance between them.

- Effective nuclear charge – Electron affinity value of the element increase as the effective nuclear charge on the atomic nucleus increases because electrostatic force of attraction between the electron being added and the nucleus increases. As the electrostatic force of attraction increases, amount of energy released is more.

- Screening or Shielding effect – Electron affinity value of the elements decreases with the increasing shielding or screening effect. The shielding effect between the outer electrons and the nucleus increases as the number of electrons increases in the inner shells.

Stability of half filled and completely filled orbitals – The stability of half filled and completely filled degenerate orbitals of a sub shell is comparatively more, so it is difficult to add electron in such orbitals and lesser energy is released on addition of electron hence the electron affinity value will decrease.



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Electron Affinities (kJ mol^{-1}), $M_{(g)}$ + e^{-} ® $M_{(g)}^{-}$ + energy

Group 1	2	13	14	15	16	17	18
H							He
72							< 0
Li	Be	B	C	N	0	F	
57	< 0	27	122	< 0	142	333	
Na	Mg	Al	Si	P	S	Cl	Ne
53	< 0	44	134	72	200	349	< 0
K	Ca	Ga	Ge	As	Se	Br	Ar
48	< 0	29	116	77	195	324	< 0
Cs				Bi	Te	1	Xe
43				106	190	295	0

4) Electronegativity:

Electronegativity of elements increases along the period while decreases down the group.

Electronegativity scales

Some arbitrary scales for the quantitative measurement of electronegativities are as under

- **Pauling's scale** –If x_A and x_B are the electronegativities of atoms A and B respectivey then
  $0.208\sqrt{\Delta_{AB}} = x_A - x_B$ if $x_A > x_B$
  or $\Delta_{AB} = 23.06\,(x_A - x_B)^2$
  $\Delta_{AB} = E_{A\text{-}B(experimental)} - E_{A\text{-}B(theoretical)}$

where E_{A-B} is the energy of A-B bond.
In a purely covalent molecule, AB, the experimental and theoretical values of bond energy A-B are equal.
So $\Delta_{AB} = 0$
or $0=23.06\,(x_A - x_B)^2$
or $x_A = x_B$
In an ionic molecule AB, $E_{A-B(experimental)}$ is more than $E_{A-B(Theoretical)}$.
Pauling assumed the electronegativity value of fluorine to be 4 and calculated the electronegativity values of other elements from this value.

- **Mulliken's electronegativity:** Electronegativitiy = (Electron Afffinity - Ionization Potential)/2
  when both are expressed in electron volt

- **Alfred Rochow's electronegativity:**
  If the distance between the circumference of outermost shell and the nucleus is r and the effective nuclear charge Z_{eff} then

  $$\frac{Z_{eff}e^2}{r^2} = \frac{0.359Z_{eff}}{r^2} + 0.744$$

  $Z_{eff} = Z - \sigma$
  Z = The actual charge present on the nucleus i.e number of protons,
  σ = Shielding constant

Factors affecting the magnitude of electronegativity

- Atomic radius: As the atomic radius of the element increases the electronegativity value decreases.
- Effective nuclear charge: The electronegativity value increases as the effective nuclear charge on the atomic nucleus increases.
- Oxidation state of the atom: The electronegativity value increases as the oxidation state (i.e. the number of positive charge) of the atom increases.
- Hybridisation state of an atom in a molecule: If the s- character in the hybridisation state of the atom increases, electronegativity also increases.

Hybridisation states	s-Character	Electronegativity
sp^3	25%	2.48
sp^2	33.33%	2.75
sp	50%	3.25

5) Valency

It is the number of univalent atoms which can combine with an atom of the given element.

- Valency is given by the number of electrons in outermost shell.
- If the number of valence electrons ≤4: valency = number of valence electrons
- If the number of valence electrons >4: valency = (8 - number of valence electrons)
- Many elements exhibit variable valence (particularly transition elements and actinoids).
- Variation in a period − Increases from 1 to 4 and then decreases from 4 to zero on moving from left to right.

- Variation in a group − No change in the valency of elements on moving down a group. All elements belonging to a particular group exhibit same valency.?

6) Metallic Characterof an Element

- Non-metallic elements have strong tendency to gain electrons.
- Non-metallic character is directly related to electronegativity and metallic character is inversely related to electronegativity.
- Across a period, electronegativity increases. Hence, non-metallic character increases (and metallic character decreases).
- Down a group, electronegativity decreases. Hence, non-metallic character decreases (and metallic character increases).