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Periodic Properties

* Periodic function:

- Atomic radius (size of atom):
The dist b/w centre of Nucleus and Outermost Orbit of an atom is known as atomic radius.
e.g. Atomic radius of some atoms are -
H - 0.37 A°
Li° - 1.73 A°
N = 1.57 A°

Atomic radius depends upon following two factors.

1) No. of Principle orbitals:
Greater the no. of principle orbital, Larger will be the Atomic radius.

2) Nuclear charge:
Higher the nuclear charge, Lower will be the atomic radius.

Note: The effect of nuclear charge is Considered only, when no. of orbitals are same.

on the basis of above factors, ~~Variation~~ Variation of atomic radius in periodic table may be summarised as -

In a period.

IA	IIA	IIIA	IVA	VA	VIA	VIIA
3Li	4Be	5B	6C	7N	8O	9F
(2,1)	(2,2)	(2,3)	(2,4)	(2,5)	(2,6)	(2,7)

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In a period form Left to right, atomic radius decrease because nuclear charge increase in no. of orbitas remain same

In a group : IA

$_3Li$ → (2,1) (n=1)

$_{11}Na$ → (2,8,1) (n=2)

$_{19}K$ → (2,8,8,1) (n=3)

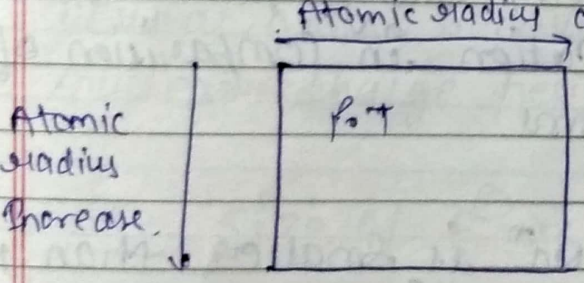
$_{37}Rb$ → (2,8,18,8,1) (n=4)

$_{55}Cs$ → 2,8,18,18,8,1 (n=5)

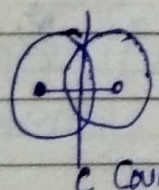
$_{87}Fr$ → 2,8,18,32,18,8,1 (n=6)

In a group from top to bottom atomic radius increase because no. of orbits increases.

Thus variation of atomic radius in periodic table may be shown as-



Note: In case of di-atomic molecules (H_2, F_2, Cl_2 etc) covalent radius is used in place of atomic radius



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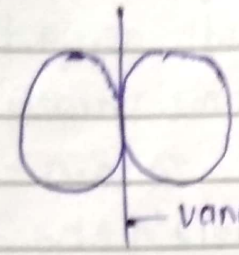
In case of mono atomic gases (He, Ne, Ar, Kr, Rn, Xe etc) Vanderwalls radius is considered in place of atomic radius.



* value of Vanderwalls radius is quite higher than covalent radius.

* Ionic radius:
A positive ion (cation) is formed by removal of e^-s from neutral atoms. The size of cation is always smaller than corresponding neutral atom because

a) Cation may have lesser no. of orbits.
b) effective nuclear charge per electron is higher in cation in comparision of neutral atoms.

Ex- Size of Na^+ is smaller than Na atom it may be explain as.

Na	Na^+
2, 8, 1	2, 8

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Size of Mg^+ is smaller than Mg

Mg & Mg^+ , Mg^{++}

2,8,2 2,8,1 2,8

In case of Mg^+, Effective Nuclear charge is higher than Mg atom. therefore size of Mg^+ will be Lower than Mg atom

As the no. of positive charge increase size decrease. eg-

$Mg > Mg^+ > Mg^{++}$

Similarly $Al > Al^+ > Al^{++} > Al^{+++}$

- A Negative ion (Anion) is formed by addition of electron in nuclear atom. The size of negative ion is always Larger than the size of Corresponding neutral atom, because in anion effective nuclear charge per electron decreases

eg- Size of $F^{\ominus}$ is larger than F atom.

F F^-

2,7 & 2,8

As the no. of negative charge increase size of ion increase.

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eg.
$O < O^- < O^{--}$

similarly
$N < N^- < N^{--} < N^{---}$

Q. Arrange the following in decreasing order of atomic radius.

1) O, S (2) Cl, Br (3) K & Ca,
4) $F^\ominus$ & Na^+ 5) Na^+ & N.

(1) O & S

$_8O = 2, 6$ $S > O$

$_{16}S = 2, 8, 6$

(2) Cl & Br

$_{17}Cl = 2, 8, 7$ $Br > Cl$

$_{35}Br = 2, 8, 18, 7$

(3) K & Ca

$_{19}K = 2, 8, 8, 1$ $K > Ca$

$_{20}Ca = 2, 8, 8, 2$

4) $F^\ominus$ & Na^+

$_9F^- = 2, 8$ $F^- > Na^+$

$_{11}Na^+ = 2, 8$

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5) Na^+ & N

$_{11}Na^+ = 2, 8$ $\quad N > Na^+$

$_7N = 2, 5$

Q Arrange following decreasing order of their radius.

K, F, Na, Cs, Br

$_{19}K = 2, 8, 8, 1$

$_9F = 2, 7$

$_{16}S = 2, 8, 6$

$_{11}Na = 2, 8, 1$

$_{55}Cs = 2, 8, 18, 18, 8, 1$

$_{35}Br = 2, 8, 18, 7$

* Ionisation Energy:-

The amount of energy required to remove one electron from outermost orbit of a neutral atom in ground state, is known as ionisation Energy. or first ionisation energy.

eg. Ionisation Energy of Li^0 is 518.0 KJ/mol

In the form of eq. 1st Ionisation Energy (E) may be represented as.

$$A_{(g)} + E \longrightarrow A^+_{(g)} + 1e^\ominus$$

If one more $e^{\ominus}$ is removed from $A^{+}_{(g)}$ then energy required will be known as second ionisation energy (I_2).

It may be shown as -

$$A^{+}_{(g)} + I_2 \longrightarrow A^{++}_{(g)} + 1e^{\ominus}$$

Similarly, IIIrd ionisation energy (I_3) may be shown as -

$$A^{++}_{(g)} + I_3 \longrightarrow A^{+++}_{(g)} + 1e^{\ominus}$$

Ionisation Energy depends upon following factors -

(1) Size of atom: Larger the size of atom, larger will be the value of electron affinity.

(2) Nuclear charge: Higher the Nuclear charge, greater will be the value of electron affinity.

3) Nature of electronic Configuration:
If electronic configuration of given atom become more stable by addition of one electron then energy will be released and if the electronic configuration of given atom becomes less stable by addition.

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of one electron then energy will be absorbed.

One the basic of above factors, variation of electron affinity in periodic table may be described as -

In a period :

IA	II A	III	IV A	V A	VI A	VII A	Zero
Li_3	$_4Be$	$_5B$	$_6C$	$_7N$	$_8O$	$_9F$	$_{10}Ne$
57 KJ	+66	-75	121	31	-142	-333	+95 KJ/m.

Here negative indicator energy released and positive sign indicator energy observed.

These values of electron affinity may be explained as -

$$_3Li^{\circ} - 1s^2, 2s^1$$

By addition of $1e^{\ominus}$, configuration of Li° becomes more stable and energy will be released.

$$_4Be - 1s^2, 2s^2$$

which one electron is added, Config. of Be become less stable and therefore energy will be absorbed.

By addition of one electron conf. of C become more stable. (half-filled p-orbitals) therefore, energy will be released

$_9F - 1s^2, 2s^2, 2p^5$

when 1 electron is added config. of F become stable therefore large amount of energy will be released

from above values, it is clear that $e^{\ominus}$ affinity mainly depends upon electronic configuration of given atom.

In general, electron affinity increases from left to right in a period with some exceptions.

In a group:

VII A

$_9F$ = −330 KJ/mol

$_{17}Cl$ = −348 KJ/mol

$_{35}Br$ = −324 KJ/mol

$_{53}I$ = −295 KJ/mole

In a group from top to bottom, electron affinity decreases with some exceptions.

eg- electron affinity of Cl is higher than F. It is due to fact that, in Cl, electron is added in 3p-orbitals, size of 3p-Orbitals is quite larger than 2p Orbital, therefore electrons of 3p will have lower repulsion than 2p-electrons, Lower the repulsion, greater will be the stability, thus $Cl^{\ominus}$ will be more stable than $F^{\ominus}$ and therefore electron affinity of Cl is higher than that of F.

$e^{\ominus}$ add. energy : ↓

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Q. First electron affinity of O atom is exothermic and 2nd electron affinity is endothermic. Explain.

Ans 1st and 2nd electron affinity of O atom may be represented as -

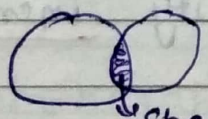
$$O_{(g)} + 1e^{\ominus} \longrightarrow O^{\ominus}_{(g)} + 142\ kJ$$

$$O^{\ominus}_{(g)} + 1e^{\ominus} \longrightarrow O^{2\ominus}_{(g)} - 844.4\ kJ$$

Second electron affinity is endothermic because there will be repulsion b/w $O^{\ominus}$ and $e^{\ominus}$ due to similar charges. Due to repulsion, stability decreases and energy is absorbed.

* Electronegativity :

A covalent bond is formed by sharing of equal no. of electrons b/w two atoms as



Shared electron pair

The shared $e^{\ominus}$ pair will be attracted by nuclei of both atom.

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The tendency of an atom to attract shared electron pair towards itself is known as electronegativity

Electronegativity depends upon following factors.

(1) Size of atom : Larger the size of atom, lower will be the value of electronegativity.

2) Nuclear charge : Higher the nuclear charge greater will be the electronegativity

on the basis of above factors, variation of electronegativity in periodic table may be shown as

In a period :

from left to right in a period, size of atoms decreases therefore. electronegativity increases from left to right.

In a group : from top to bottom in a group, size of atom increases therefore value of electronegativity decreases

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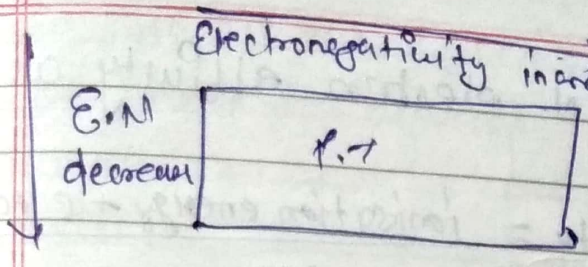

The reverse of electronegativity is electropositive nature (the tendency to loose electrons). The elements which have very high values of electronegativity will have very low electropositive nature.

The elements of zero group (He, Ne, Ar, Kr, Xe, Rn) have stable electronic Configuration therefore, they have tendency neither to loose no to gain electrons. Their electronegativity is taken as 0.0 most electronegative elements in periodic table is F and most electropositive element is Cs.

The electronegativity values of different atoms have been calculated by several scientists Pauling determined electronegativity by using bond energies while Mulliken calculated electronegetivity by using ionisation

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Energies and electron affinity as.

$$\text{Electronegativity} = \frac{\text{ionisation energy} + \text{electron affinity}}{544}$$

Here Ionisation Energy and electron affinity are in kJ/mol. if ionisation energy and electron affinity are used in electron volts then -

$$\text{Electronegativity} = \frac{\text{Ionisation Energy} + \text{electron affinity}}{2}$$

The electronegativity of some elements are

F(4.0)	Cl	Br	I	O	N	C
	3.0	2.8	2.2	3.5	3.0	2.5
H	Li	Na	Cs	etc		
2.1	1.0	0.9	0.8			

* effective nuclear charge:

The ENC may be defined as the actual nuclear charge (Z) minus the Screening effect caused by the electrons intervening b/w the nucleus and valence e⊖

ENC $Z^* = Z - \sigma$ ← shielding effect

Z = atom no.

σ = shielding or screening Const

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To calculate the Effective Nuclear charge (Z^*) we need the value of screening Const (σ) which can be calculated by using following Rules

1) Write down the electronic configuration of the element as shown below

~~(1S), (2S), (2P)~~

(1S) (2S, 2P) (3S, 3P) (3d), (4S, 4P) (4d) (4f) (5S, 5P), (5d) so on

(ii) fill the electron acc. to aufbau principle

- any electron to the right of the electron of interest do not contribute to shielding Const

- The shielding Const for each group is formed as the sum of the following contribution

→ all other electrons in the same group as the $e^\ominus$ of interest shield to an extent of 0.35 nuclear charge units except 1S group, where the other electrons contributes only 0.30

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If the group is of the [S, P] type and amount of 0.85 from each electron (n-1) shell and an amount of 1.00 for each e⁻ from (n-2) and lower shell

If the group is of the (d) or (f) + then an amount of 1.00 each e⁻ from all lying left to right to that orbital

Ex:-

Calculate ENC (Z^*) of N for 2p e⁻

$_7N = (1s^2), (2s^2, 2p^3)$

Screening const = $(\sigma) = 0.35(4) + 0.85(2)$

$= 1.40 + 1.70$

$= 3.10$

eff. N. C = $Z^* = Z - \sigma$

$= 3.90$

* Co-ordination Number:

The molecules or ion which can donate lone pair of e⁻ can exp

for Ex- NH_3, $H_2N-\ddot{O}-H$, CH_2, etc can act as ligand.

A ligent which can link with central atom through one point (atom) is known as mono dentate or unidentate ligand

NH_3, H_2O, $CN^\ominus$ etc

Polarisibility ⇒ ionic bond and covalent bond

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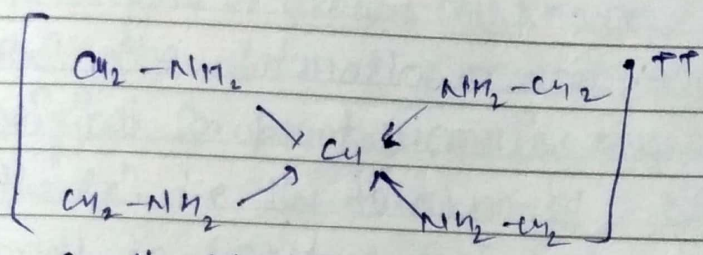
The ligand which can linked with central atom through two point (two atoms) are known as bidentate ligand.

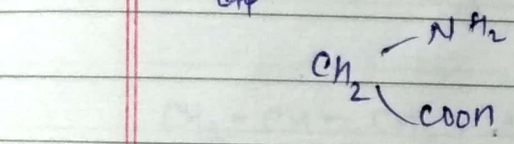
Ex: CH_2-NH_2
CH_2-NH_2

The ligand can form complex as



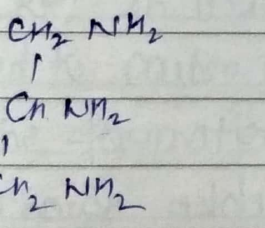
Coordination No. of Cu is $2(2) = 4$

Some other ex of bidentate ligands are



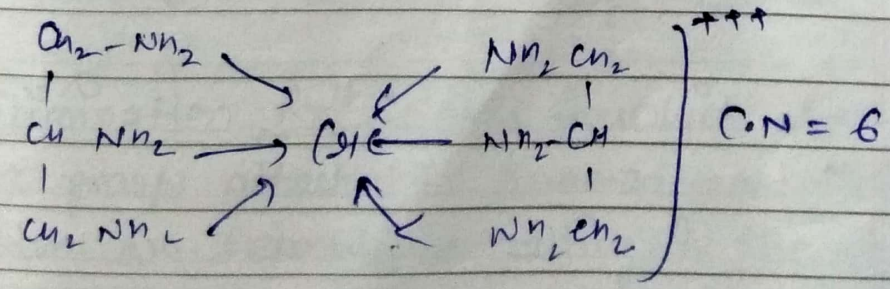
Glycine

The legend which can link with central atom through 3 point are known as ~~bigantate~~ tridentate legand



1,2,3, triamino

The legand can form complexes as



C.N = 6

Geometry depend central metal

CFSE theory

H_2O NH_3 → VSERT

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units

A ligand which can link with central metal through four point is known as tetradentate ligand

similarly the ligand which can link with central atom through 6 points are known as hexadentate ligand