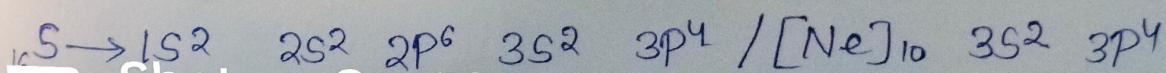
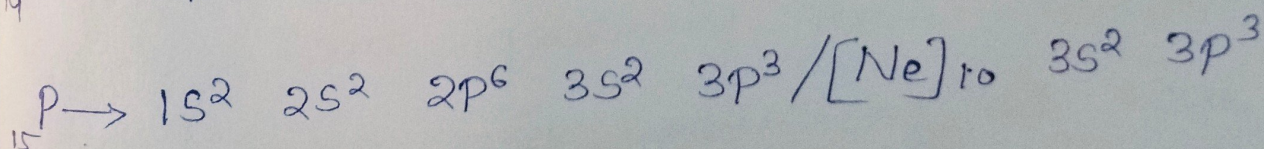
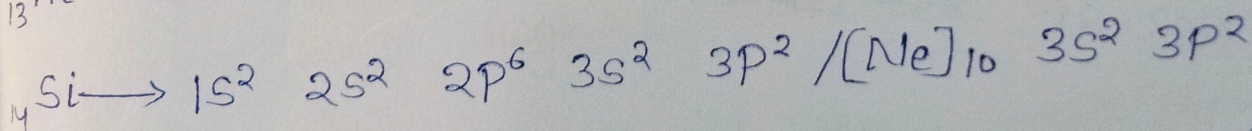
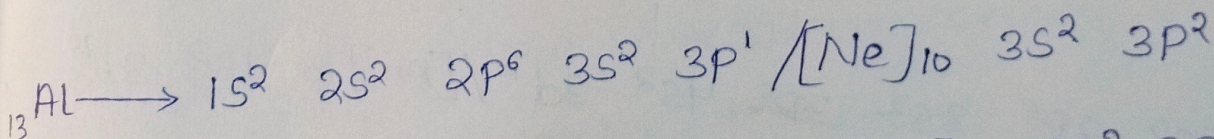
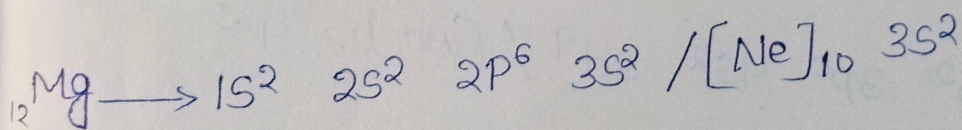
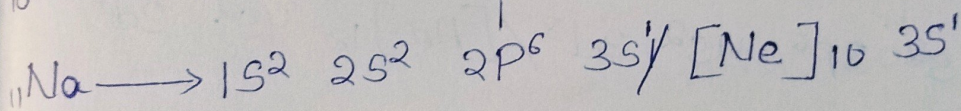
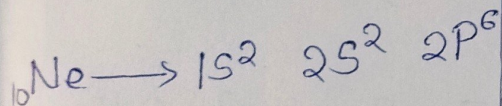
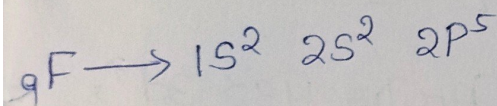
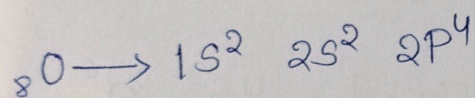
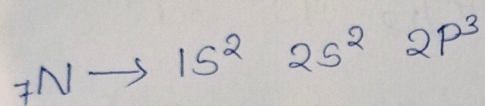
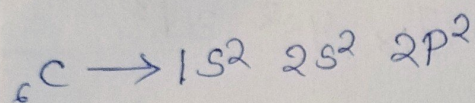
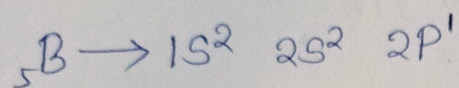
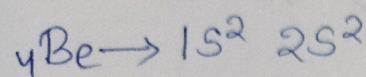
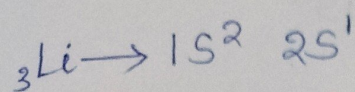
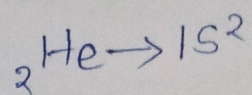
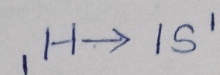


Electronic Configuration



Periodicity of properties

→ Recurrence of similar properties after certain regular interval is called periodicity of properties.

→ Periodic properties are -

① Atomic radius

② Nuclear charge

③ Ionic radius

④ Ionisation Enthalpy

⑤ Electron gain Enthalpy

⑥ Electronegativity

- $_{17}\text{Cl} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^5 / [\text{Ne}]_{10} 3s^2 3p^5$
 $_{18}\text{Ar} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 / [\text{Ne}]_{10} 3s^2 3p^6$
 $_{19}\text{K} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 / [\text{Ar}]_{18} 4s^1$
 $_{20}\text{Ca} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 / [\text{Ar}]_{18} 4s^2$
 $_{21}\text{Sc} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1 / [\text{Ar}]_{18} 4s^2 3d^1$
 $_{22}\text{Ti} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2 / [\text{Ar}]_{18} 4s^2 3d^2$
 $_{23}\text{V} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3 / [\text{Ar}]_{18} 4s^2 3d^3$
 $_{24}\text{Cr} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4 / [\text{Ar}]_{18} 4s^2 3d^4$
 $_{25}\text{Mn} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5 / [\text{Ar}]_{18} 4s^2 3d^5$
 $_{26}\text{Fe} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6 / [\text{Ar}]_{18} 4s^2 3d^6$
 $_{27}\text{Co} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7 / [\text{Ar}]_{18} 4s^2 3d^7$
 $_{28}\text{Ni} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8 / [\text{Ar}]_{18} 4s^2 3d^8$
 $_{29}\text{Cu} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9 / [\text{Ar}]_{18} 4s^2$
 $_{30}\text{Zn} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} / [\text{Ar}]_{18} 4s^2 3d^{10}$

Modern Periodic Law:-

→ It was given by Henry Moseley.

→ The physical and Chemical properties of an element are periodic function of their atomic no.

→ The element of the periodic table are divided into 4 blocks.

i) 's' Block

ii) 'p' Block

iii) 'd' Block

iv) 'f' Block

i) 's' Block elements:-

The elements in which last e^- enters into 's' subshell are called 's' block elements.
(except Helium).

Characteristics of 's' block elements:-

→ General electronic configuration ns^{1-2} they are soft metals with low melting and boiling point.

→ They have low ionisation energy.

→ They are highly Reactive.

→ They are generally form ionic compounds.

→ They are good conductors of heat and electricity in the molten state or aqua state.

→ Group 1 elements (Alkali Metals)

Group 2 elements (Alkaline earth metals)
belong to 's' block element.

ii) 'P' block elements:-

The elements in which the last e^- enters into 'p' subshell are called 'p' block elements.

Characteristics of 'P' block elements:-

→ General electronic configuration $ns^2 np^{1-6}$

→ The compound these elements are mostly co-valent in nature.

→ They include both metals and non-metals.

→ They have high ionisation energy.

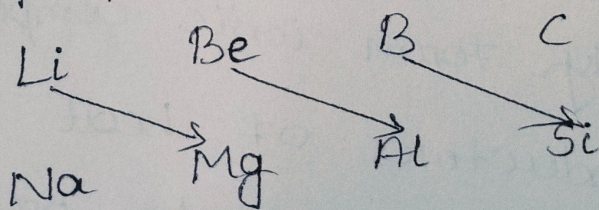
→ They show variable oxidation state.

→ They are generally bad conductors of

heat & electricity. (Except those which are metals).

→ Grp-13 to Grp-18 belong's to 'p' block element

*> Elements of 2nd period are known as bridge elements. They resemble 3rd period element in many properties.



Reason:- This is due to almost similar atomic Radius.

iii) 'd' block elements:-

The element in which the last e^- enters into 'd' subshell are called 'd' block elements.

Characteristics of 'd' block elements:-

→ General electronic configuration $(n-1)d^{1-10} ns^{1-2}$.

→ They are all metals with high melting point and boiling points.

→ The compounds of these elements are generally paramagnetic in Nature. (Presence of unpaired e^-).

→ They mostly form coloured ions and exhibit variable oxidation state.

→ They are used as catalysts and form complexes.

→ Zn, Cd & Hg have electronic configuration $(n-1)d^{1-10} ns^{1-2}$ but they don't show the properties of transition elements. So they are not considered as true transition elements.

iv) 'f' block elements:-

The element in which last e^- enters into 'f' subshell are called 'f' block elements.

Characteristics of 'f' block elements:-

→ General electronic configuration

$$(n-2)f^{1-14} (n-1)d^{0-1} ns^2$$

- These are the elements in which last e^- enters f subshell.
- They are known as inner transition elements because in the trans elements of d-block, the e^- are filled into $(n-1)d$ subshell while in inner trans elements f block the filling of e^- takes place in $(n-2)f$ subshell which happens to be an inner subshell.
- The two rows of elements at the bottom of periodic table called Lanthanides (Co to Lu) & Actinides (Th-Lr).
- They are all heavy metals. within each series the properties of the elements are quite similar.
- Most of the actinide elements are Radio-active.
- They show almost similar chemical properties and form coloured ions.

S-block	P-block	d-block	f-block
→ Last e^- enters into S-subshell → electronic configuration ns^{1-2} → low ionisation energy → high reactivity → Generally forms ionic compound. → Good conductivity of heat & electricity in aqueous or molten state. → Gr-1 element (alkali metal) → Gr-2 (alkali earth metal)	→ Last e^- enters into P-subshell → electronic configuration $ns^2 np^{1-6}$ → high ionisation energy → Variable oxidation state → Include both metal and non-metal → Covalent in nature → Bad conductivity of heat & electricity → Gr-13 to Gr-18	→ Last e^- enters into d-subshell → electronic configuration $(n-1)d^{1-10} ns^{1-2}$ → All metals with high melting & boiling point. → Paramagnetic in nature. → Formed coloured ion & exhibit variable oxidation state. → used as catalyst and form complex. → Zn, Cd & Hg have the same electronic configuration but they are no trans elements.	→ Last e^- enters into f-subshell → electronic configuration $(n-2)f^{1-14} (n-1)d^{0-1} ns^2$ → Known as inner transition elements → Two rows of elements at the bottom of periodic table called Lanthanides (Co to Lu) & actinides (Th-Lr) → heavy metals & similar elements. → actinide elements are radio-active. → form coloured ions

Periods:-

- The horizontal rows from left to right in the periodic table are called periods.
- There are seven periods in all.
- The number of elements in each period correspond to the number of electrons in different major energy levels.
- First, second and third periods are called short periods. These contain 2, 8 and 8 elements respectively.
- Fourth, fifth, sixth and seventh periods contains 18, 18, 32 and 32 elements respectively. These are called long periods.
- First period contains only two elements. This period corresponds to first main energy level whose capacity is only of two electrons.
- Second period contains eight elements. It corresponds to the second main energy level ($2s^4 2p^6$) whose capacity is of eight electrons.
- Third period contains eight elements. It corresponds to third main energy level ($3s^2 3p^6 3d^{10}$).
- It is clear from energy level diagram for multi-electron atoms that 3d-orbitals are higher in energy than 4s orbital.
- 3d orbitals are filled after filling 4s-orbital.
- Fourth period contains eighteen elements. This period corresponds to fourth main energy level. It starts with the element which receives electrons in 4s-orbital.

- After filling 4s orbital, the filling of 3d orbitals takes place.
- Fifth period also contains 18 elements (37 Rb to 54 Xe) like the fourth period.
- Sixth period contains 32 elements (55 Cs to 86 Rn) due to the filling of 6s, 4f, 5d & 6p orbitals.
- The first three elements resemble the corresponding three elements of the 5th period.
- The next 14 elements known as lanthanides are very much similar in their properties.
- Seventh period contains 32 elements (87 Fr to 118 Og). These elements are radioactive.
- Up to 92 U, the elements are naturally occurring and the remaining elements are artificially prepared.
- Hence, the elements having atomic number higher than 92 are known as synthetic elements.

Nuclear Charge:-

- Nuclear charge increases across a period and down a group.
- Nuclear charge is the positive charge of nucleus.
- It is represented by Z . Protons and neutrons are combined called nucleus.

Shielding Effect:-

The reduction of force of attraction of the nucleus on the valency electrons due to blocking effect of inner shell e^- .

Order of shielding effect:-

's' orbital > 'p' orbital > 'd' orbital > 'f' orbital.

→ 'f'-orbital have pure shielding effect.

*Effective nuclear charge = $\frac{\text{Actual nuclear charge} - \text{Shielding constant}}{\text{[No of proton present]}}$

Calculation of Shielding Constant (s)

[Slater's rules]

i) Write the electronic configuration of the element and group of the electrons in the order:

(1s), (2s, 2p), (3s, 3p), (3d), (4s, 4p), (4d), (4f), (5s, 5p), (5d, 5f), (6s, 6p) etc.

(ii) Electrons present in a group (which are on the right to the (ns, np) group) do not contribute to the screening constant.

(iii) All of the order electrons in the (ns, np) group (i.e., one electron less than the number of group electrons) shield the valence electron to an extent of 0.35 each. For an electron in

from other single electron in (s-orbital), there will be a contribution of 0.30.

iv) All the electrons in the $(n-1)$ th shell shield the valence electron to an extent of 0.85 each.

(v) All the electrons in the $(n-2)$ th or lower shell shield the valence electron to an extent of 1.00 each.

Rules for calculating ' σ ' when electron being

Shielding is in nd or nf group are:

(a) Rules (i) (ii) & (iii) are same but rules (iv) & (v) become:

(vi) All the electrons present in groups 'which are on the left of the nd or nf group' shield the valence electron to an extent of 1.00 each.

Application of effective Nuclear Charge:-

→ Cation is always smaller in size than its Parent atom.

→ This is because cation is formed by loss of e^- .

→ After loss of e^- , the effective nuclear charge of the nucleus on the valence e^- increases and thus it brings the outermost shell closer towards itself.

→ An anion is larger in size than its parent atom.

→ This is because an anion is formed by gain of e^- .

→ As the no. of electron increases the nuclear charge decreases and hence the size increases.

→ It justifies Aufbau principle according to which e^- are filled in various orbitals in the order of their increasing energy.

→ It explains that atomic and ionic size of elements of 's' and 'p' block decrease along a period.

→ This is because the effective nuclear charge increases across a period and size decreases.

Atomic radius:-

→ It is the distance between centre of nucleus to the outermost electron of the atom.

→ It is of 3 types.

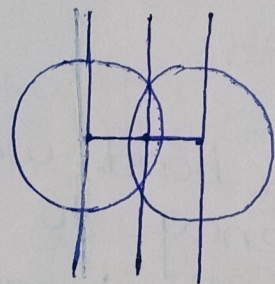
① Co-valent radius

② Vanderwaal's radius

③ Metallic radius

① Co-Valent Radius:-

→ It is defined as half of the internuclear distance between two covalently bonded atom.



Q:- If internuclear distance betⁿ two covalently bonded hydrogen atom is $0.74 \text{ \AA}$. Calculate the co-valent radius?

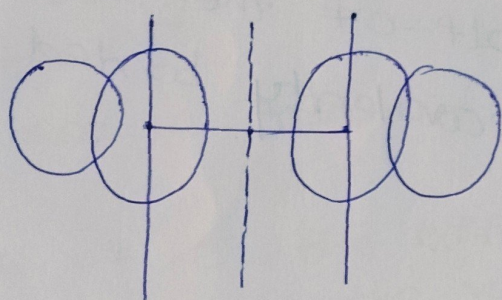
A:- Co-valent radius = $\frac{0.74}{2} = 0.37 \text{ \AA}$

② Vanderwaal's radius:-

→ It is defined as half of the internuclear distance between ^{two} non-bonded atom belonging to two neighbouring molecules in

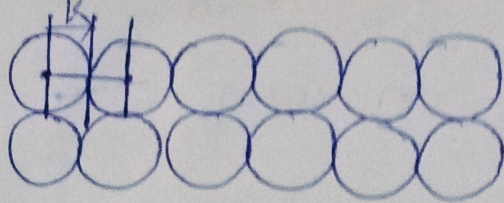
Ans

- It is mostly use for noble gases. Because these gases don't form any bond.
- Vanderwaal's radius is always greater than co-valent radius.
- Co-valent bond is form by overlapping of atomic orbital.
- This overlapping region is common to the atom which is responsible for shortening the radius. On the other hand vanderwaal's radius is calculated for bonded atom. Where no overlapping of orbital take place.



③ Metallic radius:-

- Metals don't form any bond with each other but they have the tendency to attract the atom of same elements because of presence of electron in the outer most shell.
- It is defined as half of the internuclear distance between two adjacent metal atom in metallic crystal or lattice.



Variation of atomic radius in a group.

→ Atomic radius increases in a group from top to bottom due to increase in the no. of shells.

As the distance of outermost electron from the nucleus increases therefore the force of attraction of nucleus on outermost electron decreases. Hence size of atom increases.

^{Ess} * Gallium has smaller atomic radius than aluminium.

Variation of atomic radius in a period

→ In a period, atomic radius decreases from left to right. It is because in a period no. of shells remains same but the effective nuclear charge, decreases, with increase in atomic number. and force of attraction between nucleus and outermost electron increases and atomic radius decreases.

^{Ess} * At radius of noble gases are largest in respective period because at radius of noble gases can be measured in terms of van der Waals radius which is greater than covalent radius.

Ionic radius:-

→ These are two types of ions

(1) Cations

(2) Anions.

→ Cationic radius is always less than its Parent atom.

→ Because cation has less no. of electron than the parent atom. But its nuclear charge remain same.

→ So there is a greater force of attraction of nucleus on the outermost electron and the radius of anionic decrease.

→ Anionic radius is always greater than its Parent atom.

→ Anion has greater number of electron than the parent atom. So that effective nuclear charge on the valence electron in an anion is less than the parent atom.

→ Hence the electronic cloud in an anion extends to increase the radius.

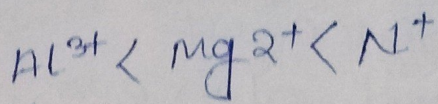
Iso electronic species:-

→ The species which contain same no. of e^- is called isoelectronic species.

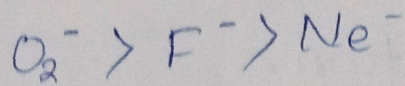
(i) Greater the +ve charge greater the radius.

(ii) Greater the +ve charge smaller the radius.

* N^+ , Mg^{2+} , Al^{3+} } Iso-electronic
 $10e^-$ $10e^-$ $10e^-$



* Ne^- , O_2^- , F^-
 $10e^-$ $10e^-$ $10e^-$



Ionisation Energy / Ionisation potential

The minimum amount of energy required to remove the valence e^- from a neutral isolated gaseous atom to convert into positively charged gaseous ion. It is called ionisation energy or ionisation enthalpy.

Unit:- kJ/mol^{-1} or kcal/mole or ev .

Factors on which ionisation energy depend

① Size of atom:-

→ Greater the size of an atom, lesser will be the force of attraction of nucleus of its valency e^- .

→ Thus, less energy will be required to remove an e^- . Hence ionisation energy will be less.

② Nuclear charge:-

→ Greater the magnitude of the nuclear charge along a period, greater will be the force of attraction of the nucleus on the valence e^- .

→ Greater will be the energy required to remove the valence electron. Hence more ionisation energy will be required.

③ Screening / Shielding effect:-

→ The shells between nucleus and valence electron to reduce the force of attraction between nucleus and valence e^- .

→ Greater the shielding effect lesser will be the ionisation energy.

4) Stable electronic configuration:-

→ Atoms with stable electronic configurations have high value of first ionisation energy.

→ The electronic configuration of an atom is stable if

(i) The orbital is fully filled

Ex:- beryllium atom.

(ii) The orbital is half filled

Ex:- Nitrogen atom

(iii) If the atom or ion has inert gas configuration

Ex:- Ne, Ar, Kr, Xe.

Variation in a ionisation energy in a group.

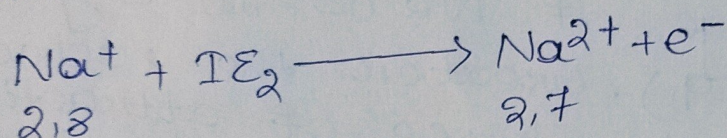
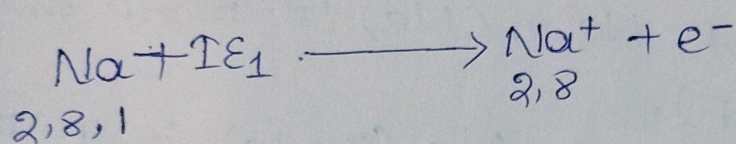
→ Ionisation energy decrease down a group from top to bottom. This is because atomic size and shielding effect increase down the group, and new shells are added. So the force of attraction of nucleus on outermost electron decrease.

Variation of ionisation energy in period

→ Ionisation energy increases across a period from left to right. This is due to the increasing nuclear charge and also electrons are added at the same shell.

Successive ionisation energy:-

The phenomenon of removing the e^- from gaseous atoms one after the other in succession is called successive ionisation energy.



$$I\text{E}_2 > I\text{E}_1$$

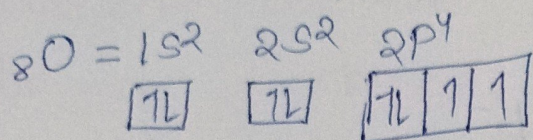
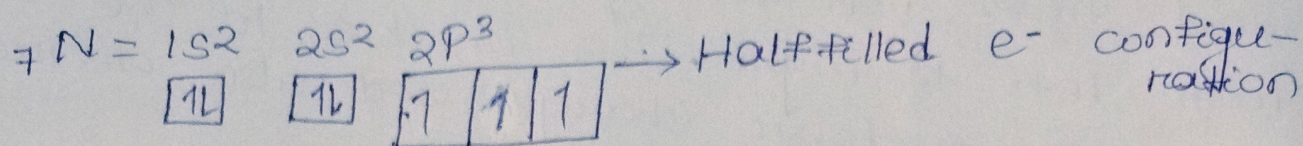
$I\text{E}_1$ = Ionisation energy required to remove the first electron.

$I\text{E}_2$ = Ionisation energy required to remove the second electron.

① Ionisation energy of Beryllium (Be) > Lithium (Li)

The nucleus charge of $\text{Be} > \text{Li}$. Greater the nuclear charge, greater will be the force of attraction betⁿ nucleus and outermost e^- . Hence 1st ionisation energy of Be is greater than Li.

② Ionisation energy of Nitrogen (N) more than oxygen (O).



→ Nitrogen has half filled configuration and half filled configuration are more stable due to symmetric and exchange energy. So, ionisation energy required more.

③ Ionisation energy of Neon (Ne) is more than that of Fluorine (F).

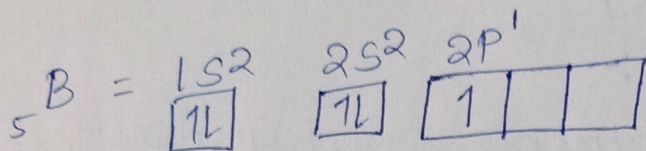
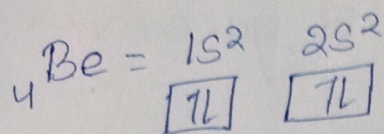
→ The nuclear charge of Ne ($Z=10$) is greater than that of F ($Z=9$). Greater the nuclear charge, greater is the force of attraction between nucleus and outermost e^- . Hence the first ionisation energy of neon is greater than that of Fluorine.

→ 2p-subshell of Ne ($1s^2 \ 2s^2 \ 2p^2x \ 2p^2y \ 2p^2z$) is fully filled. Fully filled sub-shells are more stable due to symmetry and high exchange energy. Hence, more energy is required to remove the e^- in case of Ne than in case of F. Thus ionisation energy of Ne is more than that of F.

(4) Ionisation energy of Be is more than B.

→ The nucleus charge of B greater than Be. But B has one unpaired e^- in its 2p sub-shell. But Be atom has pair e^- in 2s sub-shell. Pairing energy is not required in B.

→ s-subshell is more attracted towards the nucleus than p-subshell. Hence more energy required to remove an e^- from 2s sub-shell than 2p sub-shell. So ionisation energy of Be more than B.



Electron gain enthalpy / Electron affinity:-

→ It is defined as the amount of energy change when one e^- is added, to the outermost shell of a neutral isolated gaseous atom and convert it into a negatively charged gaseous ion is called electron gain enthalpy.

→ Unit of e^- gain enthalpy = KJ/mol or KC/mole

Successive electron affinity:-

→ The amount of energy change when electron are added successively one after another to uninegative gaseous ion is called successive electron affinity.

Factors affecting electron ^[Energy realized] affinity:-

① Size of atom:-

→ Smaller the size of atom greater will be the attraction from its nucleus for electron to be added.

→ So greater in the energy release and greater will be its electron affinity.

(2) Nuclear Charge:-

→ Greater the magnitude of nuclear charge of atom, stronger is the attraction of its nucleus for the electron to be added.

→ So greater the energy release and greater is the electron affinity.

(3) Electronic Configuration:-

→ An atom with stable electronic configuration has very little tendency to gain electron.

→ So energy has to be supplied to add an electron to form uninegative ion.

→ So there electron gain enthalpy has a positive sign. An atom has stable electronic configuration if

(i) It has half filled configuration

(ii) It has fully filled configuration

(iii) It has Noble gas configuration

Variation of electron affinity down a group:-

→ The electron affinity of elements decreases down a group due to simultaneously increase in atomic size and nuclear charge.

→ But the effect of increase in size is greater than increasing in nuclear charge.

→ As a result incoming e^- feels less attraction by the large sized atom and hence electron affinity decrease.

Variation of electron affinity across a

Period:-

→ The electron affinity usually increases across a period due to increasing effective nuclear charge, and decrease in the size of atom.

*> Electron gain enthalpy of noble gases is positive. This is because they have stable electronic configuration. Fully filled orbitals are most stable due to symmetric and exchange energy. Hence more amount of energy is required to add an electron and their electron affinity is +ve.

*> Electron affinity of Be, Mg, N and noble gases is positive:

${}^4\text{Be}(1s^2 2s^2)$ and ${}^{12}\text{Mg}(1s^2 2s^2 2p^6 3s^2)$ have fully filled s-orbital in their valence shell. Fully filled orbitals are most stable due to symmetry. Hence substantial amount of energy is absorbed to add an e^- to overcome the repulsion between negatively charged e^- being added and the negatively charged valence electrons. Their electron affinity is hence positive ($\text{Be} = +66 \text{ kJ} \cdot \text{mol}^{-1}$; $\text{Mg} = +67 \text{ kJ} \cdot \text{mol}^{-1}$).

*> Electron affinity of nitrogen is positive:

${}^7\text{N}(1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1)$ has half filled 2p-orbitals. Half filled orbitals are most stable due to symmetry. Hence substantial amount of energy is absorbed to add an electron to overcome the repulsion between negatively charged electron being added and the negatively charged valence electrons.

Its electron affinity is hence positive
($= +31 \text{ kJ mol}^{-1}$)

*) Halogens have the highest electron gain enthalpies. Halogens have the general electronic configuration of $ns^2 np^5$. Thus, these have only one electron less than the stable noble gases ($ns^2 np^6$) configuration. In order to acquire the noble gas configuration, halogens have maximum tendency to accept an additional electron & their electron gain enthalpies are, therefore high.

*) Electron gain enthalpy of Fluorine is less than that of the Chlorine

This is because of very compact size of Fluorine. It has only two shells as compared to three in Chlorine. When an electron is added to a relatively compact $2p$ -subshell, there are strong repulsions between the e^- already present and the one being added. Thus, the incoming electron does not feel much attraction. Hence, the electron gain enthalpy of Fluorine is small. On the other hand, the e^- is added to relatively large sized $3p$ -subshell in case of Chlorine which can easily accommodate the additional electron. Thus, electron gain enthalpy of Chlorine atom is large.

Calculation of Slater's rule:-

For single e⁻ species:-

H, He⁺, Li²⁺, Be³⁺

$$Z = Z_{eff}$$

For multielectronic species:-

$$Z_{eff} = Z - \sigma \text{ (value of } \sigma)$$

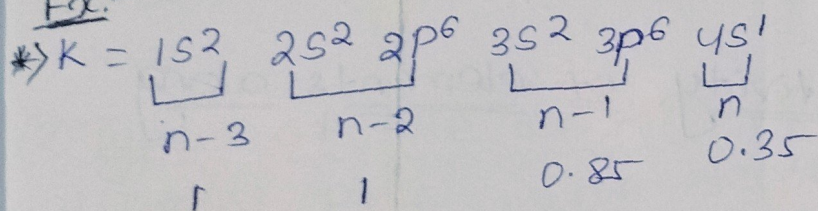
Z = actual nuclear charge

σ = Shielding constant.

Z_{eff} = effective nuclear charge

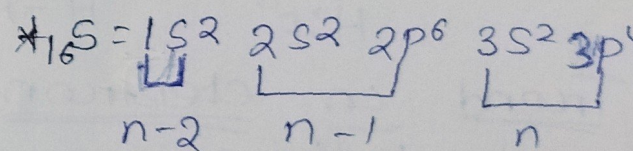
Orbital	value of σ for each e ⁻			
	(n-3)	(n-2)	(n-1)	(n)
If ns/np is outer orbital	1	1	0.85	0.35
If nd/nf is outer orbital	* 0.35 for each nd/nf e ⁻ * 1 for all remaining nd/nf e ⁻			

Ex:-



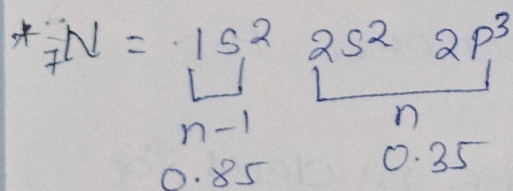
$$\sigma = 0 \times 0.35 + 8 \times 0.85 + 8 \times 1 + 2 \times 1 = 16.8$$

$$Z_{eff} = 19 - 16.8 =$$



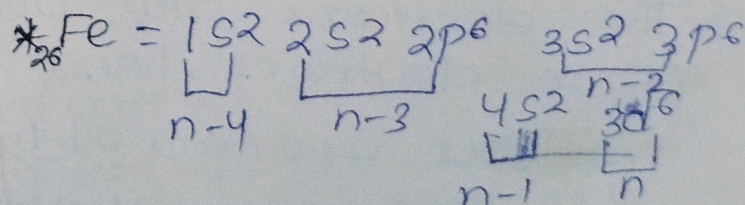
$$\sigma = 5 \times 0.35 + 4 \times 0.85 + 2 \times 1 = 10.55$$

$$Z_{eff} = 16 - 10.55 = 5.45$$



$$\sigma = 4 \times 0.35 + 2 \times 0.85 = 2.1$$

$$Z_{eff} = 7 - 2.1 = 4.9$$

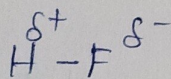
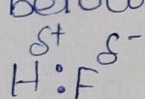


$$\sigma = 5 \times 0.35 + 2 \times 1 + 8 \times 1 + 8 \times 1 + 2 \times 1 = 21.75$$

$$Z_{eff} = 26 - 21.75 = 4.25$$

Electronegativity:-

- The tendency or power of an element in a molecule to attract the shared pair of electrons towards itself is known as its electronegativity.
- The element having higher electronegativity withdraws the shared pair of electron, more towards itself.
- Hence, it acquires some partial negative charge.
- On the other hand, the element with lower electronegativity acquires the same amount of positive charge because molecule as a whole is electrically neutral, e.g; in HF, fluorine atom acquires some partial negative charge while the hydrogen atom acquires the same amount of positive charge as shown below.



Trend in electronegativity of elements along a period:-

- As we move along a period in the periodic table, the nuclear charge goes on increasing.
- The electrons are added in the same shell. These electrons, hence do not screen the nucleus appreciably.
- The force of attraction of the nucleus for the valence electrons goes on increasing. As a result: the effective nuclear charge > screening effect.
- Hence the electronegativity of elements goes on increasing.

Trend in electronegativity of elements down a group:-

- As we move down a group in the periodic table, the nuclear charge goes on increasing.
- The electrons are added in the new shells. These electrons, hence, screen the nucleus appreciably. The force of attraction of the nucleus for the valence electrons goes on decreasing. As a result:
the effective nuclear charge < screening eff.
- Hence electronegativity of elements goes on decreasing.

Factors determining electronegativity:-

① Size of the atom:-

- The smaller size of the atom, the greater is the attraction for bonding electrons. The atoms with smaller size are more electronegative.

② Type of the ion:-

- (a) Cations are more electronegative than the atoms from which these are formed. It is because the cations are smaller in size than the corresponding atoms, e.g.; the electronegativity of Li is 1.0 while that of Fluorine atom is 4.0.
- (b) Anions are less electronegative than the atoms from which these are formed. It is because the anions are larger in size than the corresponding atoms, e.g. the electronegativity of fluoride ion (F^-) is 0.78, while that of Fluorine atom is 4.0.

③ Hybridisation:-

- The basicity of an amine depends upon the type of hybridisation of the nitrogen atom. @
- Greater the s-character of hybrid orbital, greater will be the electronegativity of N-atom.
- It lowers the donating power of electron to N-atom and hence lowers the basicity of the amine.

(4) Effect of substituents:-

- The electronegativity of an element depends upon the nature of the substituent to which it is bonded.
- The element acquires greater positive charge if the electronegativity of the substituent is higher than that of the element.

Mulliken - Jaffe's Electronegativity scales:-

- Jaffe's has defined electronegativity (χ) according to the following equation.

$$\chi = dE/dq \quad \text{--- (1)}$$

- Where E represents the total energy of the atom, related to the partial ionic charge, q by the quadratic eqⁿ.

$$E = \alpha q + \beta q^2 \quad \text{--- (2)}$$

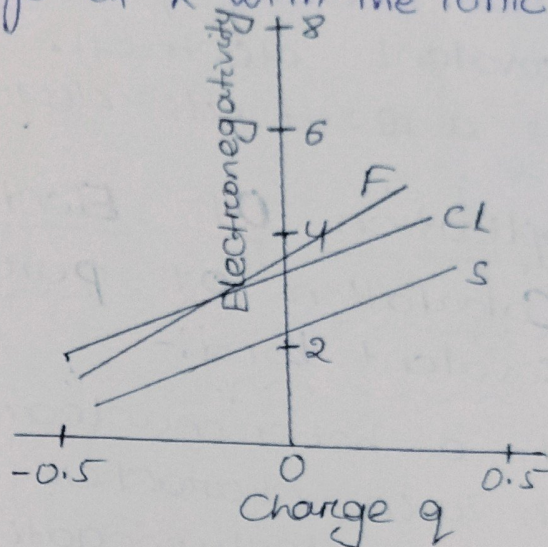
- On differentiating eqⁿ (2) with respect to 'q', we get:

$$\chi = \alpha + 2\beta q ; \chi = a + bq$$

'a' corresponds to the electronegativity of neutral atom ($q=0$). So, when $q=0$, $\chi=a$ (from eqn (3)). 'b' is called the co-efficient of electronegativity or rate of change of χ with the ionic charge q , on the atom.

(i) For small and hard atoms like fluorine (F), 'b' has large value while.

(ii) For large and soft atoms like sulphur (S) etc, 'b' has small value.



Group Electronegativities:-

→ Group electronegativities are the electronegativities of atoms, adjusted for the presence of substituents like OH, CN, COOH etc.

→ Such electronegativities are calculated from atomic electronegativities, dipole moment, kinetic data etc.

Variation of electronegativity with hybridisation and bond order:-

→ Let us consider sp , sp^2 and sp^3 hybridisation to explain the variation of electronegativity of an atom in a compound.

→ We know that an 's' electron is nearest to the nucleus due to its high penetrating power.

→ So, greater the s-character of the hybrid orbital of the central atom in a compound, greater will be its electronegativity.

Variation of electronegativity with partial charge

→ An electron is negatively charged. So, an atom, having partial positive charge will attract covalent electrons towards itself strongly. As a result, its electronegativity will be more.

Application of Electronegativity:-

1) Calculation of partial ionic character of a covalent bond:-

→ In a heteronuclear covalent molecule, A-B, an ionic character develops due to the diff. in the electronegativity of element A & B.

→ Pauling has suggested the following relation for calculating percent ionic character of a covalent bond.

$$\text{Percent ionic character} = [1 - e^{-0.25(\chi_A - \chi_B)^2}] \times 100$$

→ Where χ_A - χ_B represents the difference betⁿ the electronegativity of atoms A and B forming the covalent bond.

2) Calculation of bond length:-

→ In case of a molecule (AB) having pure covalent bond (A-B)

$$\text{Bond length (A-B)} = \text{Covalent radius of atom A in } A^0 (r_A) + \text{Covalent radius of atom B in } B^0 (r_B)$$

- When atoms A and B differ in their electronegativities, bond A-B becomes polar.
- Greater the polarity, shorter would be the bond length.

(3) Calculation of Magnitude of Bond Angles:-

- Consider a polyatomic molecule AB_2 having central atom A.
- If the electronegativity of central atom is greater than that of bonded atom B, then the bonding electron pairs would shift more towards A.
- This would result in increase in bond pair-bond pair repulsion and increase in bond angle.